" The first farmer was the first man, and all historic
nobility rests on possession and use of land."
—Emerson

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VOCATIONAL CHEMISTRY

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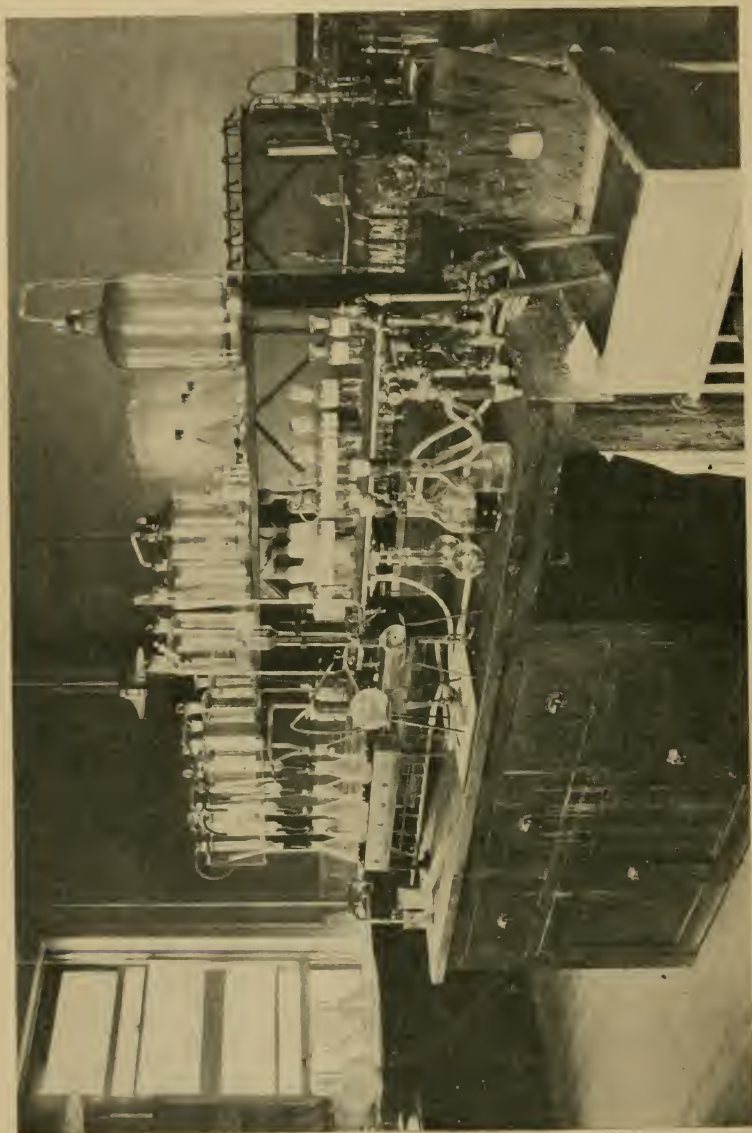
View in the chemical laboratory of a state agricultural experiment station. Here the state chemists protect the interests of the people by seeing that the analyses of feeds and fertilizers sold in the state meet the specifications. In similar laboratories dairy products, drugs, paints, water, cement, foods, and many other substances are subjected to chemical examination. Sugar factories, tanneries, dye works, soap and fertilizer factories, water works, and many other industrial concerns maintain chemical laboratories for the control of their manufacturing processes.

VOCATIONAL CHEMISTRY

FOR STUDENTS OF
AGRICULTURE AND HOME ECONOMICS

BY

JOHN J. WILLAMAN, Ph.D.

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UNIVERSITY OF MINNESOTA

70 ILLUSTRATIONS IN THE TEXT

"If vain our toil,
We ought to blame the culture, not the soil."
Pope—*Essay on Man*

QD33
W7

PREFACE

THERE are now rapidly coming into existence a large number of so-called agricultural high schools. They are located in the smaller towns or in the consolidated rural schools; they teach both agricultural and home economics courses, as the majority of their students are from the farm; they do not have large enrolments, hence mixed classes are conducted wherever possible. The latter condition has brought out the fact that the general sciences, including physics, biology, physical geography, physiology, and chemistry, are perfectly adapted to mixed classes, for the reason that these sciences are not studied as ends in themselves, but as means to an end. They serve as tools in the hands of the students for more readily grasping the subject matter of other more applied courses, such as soils, agronomy, feeding, insect control, textiles, laundering, cooking, and, indeed, most of the courses taught in such vocational high schools. They therefore serve boys and girls equally, not only in degree but in kind; and the two can use the same text in the same class, provided the text supplies information of a general nature, with no one-sided emphasis placed on applied subject matter in either agriculture or home economics.

Practically all of the present texts on elementary chemistry fall into one of two classes: those designed for students in city high schools, and those designed either for boys following agricultural courses or for girls in home economics. The present book is an attempt at meeting the above demand for an elementary chemistry text for mixed classes in the agricultural high schools. The first ten chapters are devoted to the fundamental facts and prin-

ciples of general chemistry. The other chapters are aimed
to supply the main chemical facts concerning plant and
animal growth. These facts cover the composition of the
soil; the chemical changes involved in its formation; the
maintenance of soil fertility; the compounds of the plant
and of the animal body; and the nutrition of animals.
Besides these there are chapters on the chemistry of
cooking and of cleaning, and on milk and its products.
All of these items are discussed from the viewpoint of
the chemical principles involved; working directions and
academic facts not chemical are avoided.

This selection of subject material was made so as to
give nothing but that which would be of interest and value
to both boys and girls. They are fundamentals which
should be of the common stock of knowledge of everyone;
and they should be of particular use to students of those
subjects in agriculture and home economics which have a
chemical background.

The style of presentation has, of course, been made
as simple as possible. Many of the laws and subjects that
are conventionally presented in texts on general chemistry
are omitted; among these are ionization, dissociation, the
gas laws, and the laws of multiple and reciprocal propor-
tions. The writer has found that these topics are not
essential to an understanding of the chemistry which is
needed by these students.

The text is designed for a full year's course, including
classroom and laboratory work. Where facilities are not
available for the students to perform the laboratory work,
this can be done by the instructor as demonstration. If
only one term is given to chemistry, the text alone can be
followed, with a limited amount of demonstration work.
In this case the last two or three chapters may have to be
omitted, which can be done without impairing the unity of
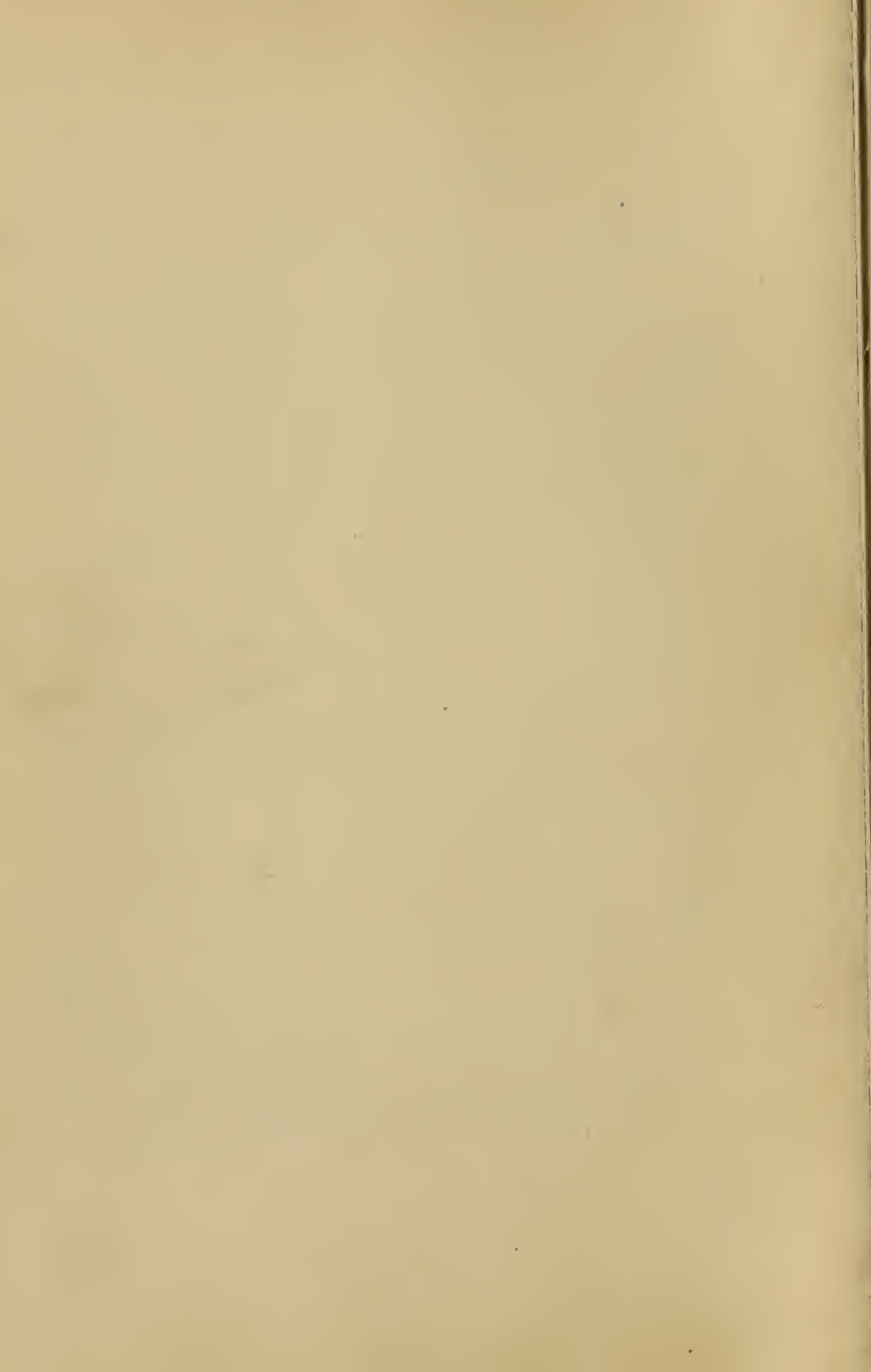the course.

Realizing that no doubt many sins both of commission and omission are present, the writer will welcome any suggestions for the improvement of this volume.

The writer wishes to acknowledge the aid of the various firms and institutions which furnished illustrative material: Union Sulfur Co., Fig. 39; Swift & Co., Figs. 15, 38, 56, 61, 64, 66, 67; Detroit Steel Products Co., Fig. 34; Harbison-Walker Refractories Co., Figs. 31, 32; American Agricultural Chemical Co., Fig. 49; Northrup, King & Co., Fig. 65; Pittsburgh Filter Mfg. Co., Figs. 6, 7; Florida Wood Products Co., Fig. 13; Acheson Graphite Co., Fig. 21; also the help of his wife in making many of the drawings, in correcting proof, and in giving valuable suggestions on points of pedagogy.

J. J. W.

St. Paul, Minnesota.
    August, 1920.

CONTENTS

VOCATIONAL CHEMISTRY

CHAPTER I

THE FIELD OF CHEMISTRY

A Changing World.—If we stop a moment to consider what is the most characteristic thing about this world in which we live, we shall no doubt say that it is the continual changes that are taking place. All of our activities involve changes. If we walk across the room, move a book, or throw a stone, we have brought about changes in location of the various objects. If we heat a piece of iron red hot, we have changed its temperature and color, and even its size. If we strike a match, we have changed the head and the splinter of wood into some fumes and a little pile of ashes. In most places on earth the temperature changes considerably from summer to winter. When we eat a meal, within a few hours the bread and meat are converted into the tissues of our body. When we snap a picture with a camera, we change the film or plate in some way that enables a picture to be developed on it. Even while we sleep remarkable changes are going on in the body. Rocks slowly but surely decay and crumble and change into soil. A tree changes the substances of the soil into a delicious apple; and if the juice of the apple is properly fermented, it is changed into vinegar. Our foods change considerably in flavor and texture during cooking. Water very frequently is changed into a solid or into a gas. Astronomers even tell us that the sun is changing—it is gradually cooling off. In short, we can say that *everything in the universe is continually changing.*

1

Physical and Chemical Changes.—It will be noticed in the above brief list of familiar changes that in some cases the change creates new substances; as, the digestion of food, the burning of the match, the growth of the apple. In other cases the substance remains the same, but is changed as regards color, temperature, size or location. The first are spoken of as chemical changes, the latter as physical changes. Thus, a match may be whittled into an infinite number of fine shavings without changing the chemical composition of the wood in any way, for the shavings would still contain the same chemical substances in the same proportions as in the original whole match. When, however, the match is ignited, the substances of the wood unite with the oxygen of the air and are converted into various other substances which constitute the smoke and the fumes and the ashes. These two cases illustrate a physical and a chemical change, respectively. The science of physics deals with physical changes, the science of chemistry with chemical changes. In other words, *chemistry is the study of the composition of substances and the changes in composition which these substances undergo.*

Branches of Chemistry.—It will be readily seen that the field of chemistry is a large one. It involves the make-up of the air, of waters, of rocks, of soils, of plants and animals, foods and clothing, fuels and oils, drugs and medicines, glass and pottery, paper, leather, building materials, manures—in short, every conceivable kind of substance. In fact, the field of chemistry is so large that it is conveniently divided into several branches, and we have agricultural chemistry, mineral chemistry, industrial chemistry, physiological chemistry.

The particular things in which we are interested, of course, are the chemical processes involved in our daily life in the home and on the farm. We want to know what happens when silage ferments, when eggs spoil, when hay

cures, when manure becomes the humus of soil; what the compounds in foods are, and what becomes of them during digestion; why some water is "soft" and some is "hard"; what the chemistry of soap-making and bluing and washing powders is; why butter becomes rancid, cement hardens, and fruit juices jell. We want to be on speaking terms with the gases of the atmosphere, the common acids and salts, the elements of the soil and of plant and animal bodies, the constituents of coal and petroleum, and many other things which are vital to our everyday lives. If we know some of these things, it will be like going into a forest and being able to call the trees by name; we shall appreciate the more this old world of ours with its manifold changes. But, most of all, we shall be better able to understand the other arts and sciences which have a chemical basis—nutrition, soil fertility, laundering, textiles, metallurgy, cooking, food manufacture and preservation. If this is kept in mind, the study of the succeeding chapters will be more profitable.

QUESTIONS

1. Arrange the following into two groups, one for physical changes and the other for chemical changes:—
>>> Burning of wood.
>>> Contraction of steel rails in winter.
>>> Fermentation of sauerkraut.
>>> Mixing sugar and salt.
>>> Mixing plaster of Paris with water.
>>> Sawing a board in two.
>>> Freezing of water.
>>> Running milk through a cream separator.
>>> Grinding grain.
>>> Boiling an egg.
>>> Making soap.
2. Explain how a knowledge of chemistry will help us to understand many facts and processes in our everyday life.
3. With what does chemistry deal?

THE COMPOSITION OF THE UNIVERSE

Chemical Elements and Compounds.—As we look about us, we realize that the world is made up of a large number of different substances—countless thousands of different substances, in fact. The soil contains a host of different materials; there are hundreds of different minerals known; there are half a dozen gases in the atmosphere; every organ and tissue of every kind of animal and plant contains different substances; our factories take the materials of nature and make hundreds of new drugs, chemicals and foods out of them; petroleum is pumped from the ground and then separated into many different substances. Apparently it is a very complex world in which we live.

Elements.—But the science of chemistry has revealed the following very remarkable fact:

There are known to exist only about eighty fundamental basic substances, out of which everything in the universe is made. That is, all of the thousands of different materials mentioned above really involve only eighty different simple substances. *These eighty fundamental substances are called the chemical elements.*

Compounds.—Various combinations of these elements, and various proportions of them, make possible the enormous number of different materials that we know to exist in the world. In fact, most of these materials are composed of but a very few of the chemical elements. Thus, common salt is made up of two elements, sodium and chlorine; cane sugar of three, carbon, hydrogen and oxy-

4

gen; water, two; fat, three; egg white, five; alcohol, three; gasoline, two; very few drugs contain more than four, and very few minerals more than seven. *When two or more chemical elements unite to form a substance of definite composition, that substance is called a chemical compound.* Thus, salt, sugar, alcohol, limestone, quinine, etc., are compounds, as they consist of two or more elements united together chemically in certain definite and fixed proportions.

What Chemistry Includes.—In the first chapter we learned that chemistry is the study of the composition of substances. From a little different viewpoint, we now see that chemistry is really a study of elements and compounds and the changes which they undergo, since every substance in the world is either an element or a compound, or a mixture of them. We can say, then, that the world is composed of eighty chemical elements and innumerable compounds of those elements. It has been the chemists' job to find out what these elements are, to separate these various compounds, to determine what elements are in them and what they can be used for, and to combine elements together to make new compounds. Thus chemical study has told us what gases are in the atmosphere, how much of each is present, and for what purpose each is there; it has told us what the mineral compounds of the soil are, and which of them the plants use; it has told us how to manufacture explosives, chloroform, hydrogen peroxide, sugar, matches, steel, electric batteries, photographic plates; it has even revealed the nature of volcanic gases, and the composition of the sun and stars.

Early Chemistry.—This is far different from the objects of chemical study three hundred years ago. At that time there were just two things sought for by chemists: a method of converting iron, lead, and other cheap metals

into gold, and a medicine which would cure all diseases. They believed there were only four elements: earth, air, fire, and water. Concerning these, we now know that they are not elements at all; that the earth contains many elements and compounds, that the air contains at least seven gaseous elements and two compounds, that water is a compound of two elements, and that fire is not a substance, but a violent chemical change.

An Element Defined.—What, then, is a chemical element? From what we have seen above, it will be apparent that *a chemical element is a substance which consists of but one thing; it cannot be divided or decomposed or resolved into any simpler substance.* For example, sulfur is an element, because so far chemists have not been able to separate it into two or more other substances; it consists of but one thing. Iron, gold, oxygen, iodine, mercury, are elements for the same reasons.

List of Elements.—In the accompanying table is given a list of the chemical elements now known. The commoner and more important ones are in black-faced type. As has been said, only some eighty chemical elements have been found; more, probably, exist, for every now and then a new one is discovered. Of these eighty elements, however, the large majority are of very rare occurrence. It can be safely said that in our ordinary daily life we never see substances which involve over thirty of these elements; only about fifteen are found in the animal body, and only ten are absolutely necessary for plant growth. Therefore, the chemistry which will interest us will involve only a few of the known chemical elements.

The Most Abundant Elements.—Furthermore, approximately 97 per cent of the mass of the earth's crust, including the water and the air, is made up of but eight elements; the other seventy-two put together make up only

about one-thirtieth of the earth. Figure 1 shows what these abundant elements are and the amount of each, in the solid earth, in the oceans, and in the air.

No doubt by now the student has recognized that many
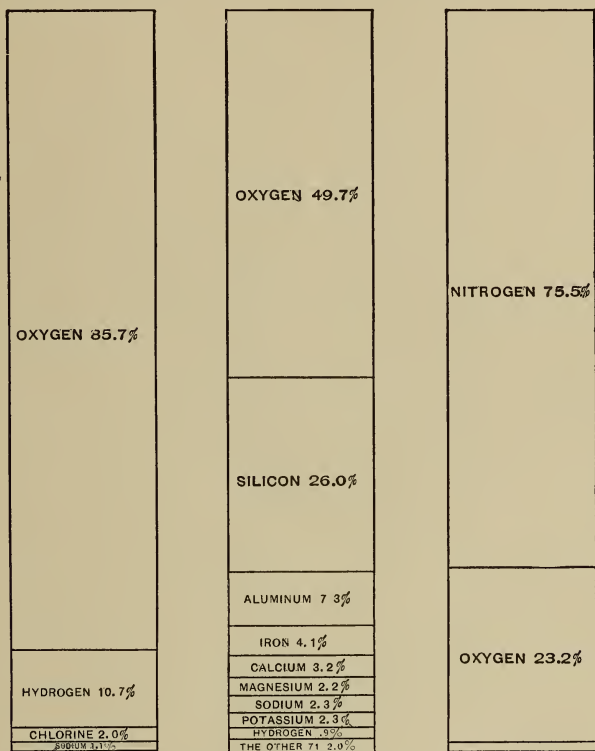
FIG. 1.—Diagram showing the comparative abundance of the chemical elements in *A*, the oceans; *B*, the air; and *C*, the whole earth. Notice the prominent place of oxygen in all three, and the scarcity of nitrogen in all but the atmosphere.

familiar substances are found in the list of chemical elements, as iron, lead, tin, oxygen, nitrogen, phosphorus, sulfur, carbon. Also, many common substances such as starch, water, sand, salt, and glass, are not in this list; therefore, they must be compounds, or mixtures of compounds.

TABLE I

The Known Chemical Elements With Their Symbols and Atomic Weights

Name	Symbol	Atomic weight	Name	Symbol	Atomic weight
Aluminum	Al	27.1	Molybdenum...	Mo	96.0
Antimony	Sb	120.2	Neodymium....	Nd	144.3
Argon	A	39.8	Neon	Ne	20.2
Arsenic	As	74.9	**Nickel**	Ni	58.6
Barium	Ba	137.3	Niton (Ra ema-		
Bismuth	Bi	208.0	nation)	Nt	222.4
Boron	B	11.0	**Nitrogen**	N	14.0
Bromine	Br	79.9	Osmium	Os	190.9
Cadmium	Cd	112.4	**Oxygen**	O	16.0
Cæsium	Cs	132.8	Palladium	Pd	106.7
Calcium	Ca	40.0	**Phosphorus**	P	31.0
Carbon	C	12.0	Platinum	Pt	195.2
Cerium	Ce	140.2	Potassium	K	39.1
Chlorine	Cl	35.4	Praseodymium..	Pr	140.9
Chromium	Cr	52.0	Radium	Ra	226.0
Cobalt	Co	58.9	Rhodium	Rh	102.9
Columbium	Cb	93.1	Rubidium	Rb	85.4
Copper	Cu	63.5	Ruthenium	Ru	101.7
Dysprosium	Dy	162.5	Samarium	Sa	150.4
Erbium	Er	167.7	Scandium	Sc	44.1
Europium	Eu	152.0	Selenium	Se	79.2
Fluorine	F	19.0	**Silicon**	Si	28.3
Gadolinium	Gd	157.3	**Silver**	Ag	107.8
Gallium	Ga	69.9	**Sodium**	Na	23.0
Germanium	Ge	72.5	Strontium	Sr	87.6
Glucinum	Gl	9.1	**Sulfur**	S	32.0
Gold	Au	197.2	Tantalum	Ta	181.5
Helium	He	4.0	Tellurium	Te	127.5
Holmium	Ho	163.5	Terbium	Tb	159.2
Hydrogen	H	1.0	Thallium	Tl	204.0
Indium	In	114.8	Thorium	Th	232.4
Iodine	I	126.9	Thulium	Tm	168.5
Iridium	Ir	193.1	**Tin**	Sn	118.7
Iron	Fe	55.8	Titanium	Ti	48.1
Krypton	Kr	82.9	Tungsten	W	184.0
Lanthanum	La	139.0	Uranium	U	238.2
Lead	Pb	207.2	Vanadium	V	51.0
Lithium	Li	6.9	Xenon	Xe	130.2
Lutecium	Lu	175.0	Ytterbium	Yb	173.5
Magnesium	Mg	24.3	Yttrium	Yt	88.7
Manganese	Mn	54.9	**Zinc**	Zn	65.3
Mercury	Hg	200.6	Zirconium	Zr	90.6

Chemical Symbols.—One of the first things the student must do is to familiarize himself with the names and the symbols of the commonest chemical elements. The symbol is the short way of writing the name of the element. It is often the first letter, sometimes it is the first two letters, sometimes it is the first letter of the Latin name of the element. Thus O is oxygen, Ca is calcium, Fe is iron (Latin *ferrum*).

Atoms.—The symbol, however, means more than an abbreviation. It stands for a single particle of that element. This single particle is called the *atom*. *It is the smallest particle of an element that can exist.* It is thousands of times smaller than can be seen by even a powerful microscope. In studying the characteristics and properties of an element, we therefore study the characteristics and properties of the atoms of that element. If an element is green, it is because each atom is green; if an element tastes sour, it is because each atom is sour. Therefore, it is well to realize from the beginning that the study of chemistry is really a study of atoms—their appearance, characteristics, uses, occurrence, and combinations— although a single atom, of course, is never handled.

Molecules.—Now, when two or more elements combine to form a compound, it must be the atoms of the elements that do the combining. And since the atom is the smallest particle of the element, these combined atoms must form the smallest particle of the compound. *This smallest possible particle of a compound is called the molecule.* Thus, when one atom of sodium, represented by Na, combines with one atom of chlorine, represented by Cl, one molecule of sodium chloride or common salt is formed. This molecule is represented by NaCl. It is the smallest particle of that substance that can exist. It has all the characteristics of a spoonful of salt, for the spoonful is white, dissolves in water, and has its characteristic taste only

because each minute molecule of the salt is white, dissolves in water, and has that taste. Just as we said that a study of an element means the study of an atom of that element, likewise the study of a compound means the study of a molecule of it.

Chemical Equations.—There is one more thing we must consider in this chapter, and that is the method of writing chemical changes. In the above example, the formation of sodium chloride, or common salt, could be written thus:

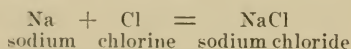<p style="text-align:center">sodium combined with chlorine forms sodium chloride</p>

or it could be written:

<p style="text-align:center">sodium + chlorine = sodium chloride.</p>
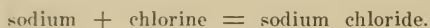
It is very much simpler, however, to write it thus:

$$Na + Cl = NaCl$$
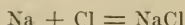
Usually the latter two forms are combined:

$$Na + Cl = NaCl$$
<p style="text-align:center">sodium   chlorine   sodium chloride</p>

Such a method of recording a chemical change is called a chemical equation. The above equation tells us that one atom of sodium combines with one atom of chlorine to form one molecule of sodium chloride. The formation of quicklime would be shown in the following equation:

$$Ca + O = CaO$$
<p style="text-align:center">calcium   oxygen   calcium oxide (quicklime)</p>
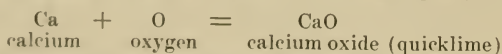
The formation of water demands a little different treatment, since two atoms of hydrogen combine with one of oxygen:

$$2H + O = H_2O$$
<p style="text-align:center">hydrogen   oxygen   hydrogen oxide (water)</p>
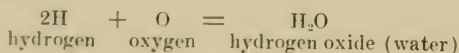
It will be noted that the 2 is written *before* the H when the letter stands alone as an element, but that it is written *after and below* the H in the molecule. In the molecule of H_2O, there are two atoms of hydrogen and one of oxygen; the figure 1, however, is not written, but is always under-

stood. Likewise in the formula for calcium oxide, CaO, there is one atom of calcium and one atom of oxygen, although the figures are not written.

Atomic Weights.—From the very beginning of modern chemistry it was recognized that, although one atom of Na combines with one atom of Cl to form NaCl, it is not one ounce of Na to one ounce of Cl; nor does H_2O represent two ounces of H combined with one of O. *It was found that all elements combine in different proportions by weight and hence that the atoms of the different elements weigh different amounts.* For example, an atom of oxygen weighs 16 times as much as an atom of hydrogen; an atom of sulfur twice as much as an atom of oxygen. By extremely careful methods of study, the relative weights of all the elements have been determined. Since hydrogen is found to have the lightest atom of all the chemical elements, its relative weight is taken as 1. And since the oxygen atom is 16 times heavier, its weight is taken as 16; and as sulfur is twice as heavy as oxygen, it must be 32 times as heavy as hydrogen, and its relative weight is taken as 32.

In like manner, every element has been given such a relative weight, and *these weights are known as the atomic weights.* They are given in Table I. The atomic weight is not the weight of the atom in ounces or grams or any other unit of weight; *it simply represents the number of times the atom of that element is heavier than the atom of hydrogen.*

Molecular Weights.—The atomic weight of sodium is 23; that of chlorine is 35.5. In the equation

$$\text{Na} + \text{Cl} = \text{NaCl}$$
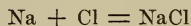
23 parts by weight of Na combine with 35.5 parts by weight of chlorine, forming 58.5 parts by weight of NaCl. It makes no difference whether we use 23 and 35.5 ounces,

or 23 and 35.5 grams, or tons; the relation still holds, and
there would result 58.5 ounces, grams, or tons of NaCl.
The number 58.5 is called the *molecular weight; it is the
sum of the atomic weights in the molecule.*

What an Equation Represents.—We are now in a posi-
tion to examine carefully the full meaning of a chemical
equation. Since all chemical changes can be written in
the form of equations, and since the science of chemistry
deals with chemical changes or reactions, it is very neces-
sary at the outset of our course thoroughly to understand
and appreciate the meaning of such equations. The ex-
amples given above are very simple; many, however, are
very complicated. For a simple example, consider the
chemical change involved in the burning of a piece of
charcoal. Charcoal is carbon, and when it burns it com-
bines with oxygen thus:

$$\text{C} \quad + \quad 2\text{O} \quad = \quad \text{CO}_2$$
carbon   oxygen   carbon dioxide (carbonic acid gas)
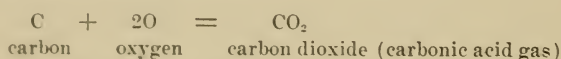
If, however, we try to represent in an equation the burn-
ing of a piece of wood, we meet with complications. The
wood is not a single substance like charcoal; it contains
dozens of different substances, each of which combines
with oxygen in its own particular way, and hence would
demand an equation all of its own. Therefore, not one
equation, but many of them would be necessary to prop-
erly express the chemical changes involved in the burning
of a piece of wood. Only one chemical change at a time,
then, can be considered in writing a chemical equation.

It will be noticed in all of the above examples that the
equation tells us what substances unite, or " react," and
what substance or substances are formed. It also gives us
the *formula* of each substance; that is, the kinds of atoms
and the number of each, in each molecule. Thus the for-
mula of carbon dioxide, commonly known as carbonic acid

gas, is CO_2, which means that the molecule of carbon dioxide consists of one atom of carbon and two atoms of oxygen. Every compound has a chemical name. The chemical name tells of what the compound is made. Thus, the chemical name for water is hydrogen oxide; the latter we get from the formula H_2O. The chemical name for common salt is sodium chloride, and this name we get from the formula NaCl. The system of chemical names will be explained later. Furthermore, an equation tells us the relative quantities of the substances involved. Thus the above equation shows that 12 grams, or ounces, or tons, of carbon combine with 2×16, or 32, grams, ounces, or tons of oxygen, to form 44 grams, ounces, or tons of carbon dioxide.
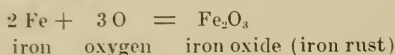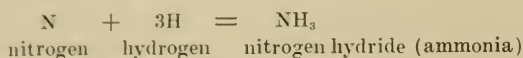
Important Meanings of an Equation.—The following points, then, should be kept in mind as regards a chemical equation:

1. A chemical equation represents a single chemical reaction.

2. The equation tells us what substances react with each other, and what products are formed.

3. It tells us how many atoms and molecules of each substance are involved.

4. It tells us the formula of each substance, and the formula enables us to give the compound a chemical name.

5. It tells us the quantities by weight of each substance involved in the reaction.

QUESTIONS

1. How does a chemist prove whether a substance is an element or a compound?
2. How many elements are there in the world? How many compounds?
3. Name ten elements and ten compounds.
4. How can you find out from this chapter which of the following are elements and which are compounds? Tin, granite, soap, Epsom salts, blue vitriol, nitrogen, silver.
5. Name some of the things that chemistry has taught us and has produced for us.

6. What is the most abundant chemical element in the air? In the ocean? In the crust of the earth?
7. What is the exact meaning of the formula of a compound?
8. What five things does each of the following equations tell us?

$$S + 2O = SO_2$$
sulfur   oxygen   sulfur dioxide

$$H + Cl = HCl$$
hydrogen   chlorine   hydrogen chloride (muriatic acid)

$$N + 3H = NH_3$$
nitrogen   hydrogen   nitrogen hydride (ammonia)

$$2Fe + 3O = Fe_2O_3$$
iron   oxygen   iron oxide (iron rust)
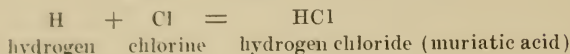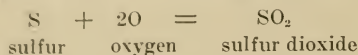
9. Why should we learn to know substances by their chemical name as well as by their common name?

LABORATORY EXPERIMENTS

1. To Study Examples of Physical Changes.—(a) Weigh out 5 g. of sugar, and dissolve in 25 c.c. of water in a porcelain evaporating dish or a beaker. Evaporate carefully to dryness, using a water bath if possible. Carefully scrape out the residue and weigh it. Has the treatment with water changed the amount of sugar? Test it as follows: Take portions of it, and portions of the original sugar and compare their taste. Examine them under the lens. Heat them in test tubes. What changes did the sugar undergo? Are these changes physical or chemical? How does dissolving in water affect sugar? What does this experiment teach us concerning the process of obtaining granulated sugar from sugar beets and sugar cane?

(b) Heat a piece of glass rod or tubing very gently at first and then very strongly. What changes has the piece of glass undergone? Mention three other examples of similar changes.

(c) Place a few flakes of iodine in a small porcelain dish, and cover the dish with a small piece of glass. Stand the dish on asbestos on a tripod or ring stand and heat very gently and slowly. What becomes of the iodine? What kind of a change has it undergone? How do you know?

2. To Study Examples of Chemical Change.—(a) Heat a little sugar in a test tube, at first gently, then strongly. What kind of a change has taken place? How do you know? What remains in the tube?

(b) Heat a similar amount of sugar in an open dish, at first gently, and then with the hottest flame. How does this change compare with that in (a)? What is the essential difference between them?

3. To Distinguish Between Chemical and Physical Changes.—(a) Test some powdered sulfur with a magnet. Test its solubility in carbon bisulfide by mixing a little sulfur with the liquid, pouring out some of the clear solution onto a watch glass, and then evaporating down in a warm place, *but away from flames!* Test a little of the sulfur in a test tube with dilute hydrochloric acid.

(b) Make the same tests on some powdered iron.

(c) With a mortar and pestle, thoroughly mix about 3 g. of sulfur and 5 g. of iron powder. Describe the changes that have taken place. Test small portions of this mixture with the carbon disulfide, with the magnet, and with hydrochloric acid. If the sulfur appears to stick to the magnet, gentle shaking will release it. Has a physical or a chemical change taken place in the iron and in the sulfur?

(d) Place the rest of the mixture from (c) in a test tube, and carefully heat until the mass begins to glow like a red-hot coal. When the action has ceased, break the tube by immersing in water, pick out the fused iron and sulfur, and perform the tests as in (a), (b) and (c). What type of change has taken place? How do you know? What substance has been formed? State all the physical changes that took place during this experiment, then all the chemical, and name the products formed during each chemical reaction. Write these reactions in the form of equations.

4. **To Study Examples of Chemical Combination.**— (To be done by instructor.) (a) In a porcelain dish, or on a piece of asbestos, place together a small piece of phosphorus (never handle with the hands) and a few flakes of iodine. What product is formed? Write the reaction in the form of an equation. Why is this reaction called combination?

(b) By means of tongs hold a piece of magnesium ribbon in the flame. Describe the result. What is the " smoke "? Write the equation. Why is this a combination change?

(c) State three other examples of chemical combination.

(d) What reactions in experiments 2 and 3 are chemical combinations?

5. **To Study Examples of Chemical Decomposition.**—(a) Heat a few crystals of copper sulfate (blue vitriol) in a porcelain dish. What change in form have the crystals undergone? What is the white substance left? Write the equation. Change this example of decomposition into one of combination by adding a few drops of water to the white residue and thus restoring the blue vitriol.

(b) What reaction in experiment 2 is decomposition?

(c) Is the burning of a match combination or decomposition? Why?

THE ATMOSPHERE

Taking up now some of the commonest elements and their compounds, so as to familiarize ourselves with the chemical nature of the world about us, we naturally first look to the atmosphere in which we live, because the atmosphere is a mixture of two very important elements, oxygen and nitrogen, and at least two important compounds, carbon dioxide and water vapor.

Nature of the Atmosphere.—Our atmosphere is a thin layer of gas surrounding the earth. It is thin in proportion to the diameter of the earth; for while the latter is 8000 miles, the air is probably not over 200 or 300 miles in depth, and the greater bulk of it by weight is within the first forty or fifty miles. In fact, we know that mountain climbers and aviators have great difficulty in breathing when they rise to even four miles. These facts show that air has weight and is pulled towards the surface of the earth by the force of gravity. The actual weight of air over every square inch of surface at sea level is about 15 pounds, or over 41,000 tons per acre. This weight of the air is very important, as it propels our windmills and sailing vessels, and forces the water up into our pumps when the valves make a partial vacuum.

The Air and Mixture of Gases.—That the air is not a single substance but a mixture can be proved in two interesting ways. The first is to allow a piece of phosphorus to burn in an enclosed volume of air, as under a bell jar * (Fig. 2). The phosphorus will unite with the oxygen of

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* See experiment 8 at end of chapter.

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the air as long as there is any present, forming phosphorus pentoxide according to the following equation:

$$\underset{\text{phosphorus}}{2P} \quad + \quad \underset{\text{oxygen}}{5\,O} \quad = \quad \underset{\text{phosphorus pentoxide}}{P_2O_5}$$
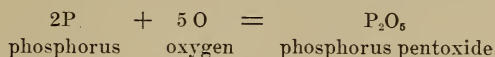
This compound is the white fumes which gradually dissolve in the water and disappear. The gas now left in the bell jar is practically all nitrogen, with traces of carbon dioxide and water vapor. It will be noticed that the water has risen in the jar to take the place of the oxygen removed. If the oxygen of the air had been combined with the nitrogen, the phosphorus could not have removed it. This proves that the oxygen, nitrogen and other gases are simply mixed together, like the mixture of gasoline vapor and air in a carburetor.

Fig. 2.—Apparatus used in demonstrating that air is a mixture of gases. The burning piece of phosphorus on the cork under the bell jar removes the oxygen from the enclosed air. This creates a partial vacuum, and the water rises in the jar. If all the oxygen were removed, the water would occupy about one-fifth of the space under the jar.

The other proof that air is a mixture is the fact that the various constituents have different densities, and the heavier ones have a tendency to settle towards the surface of the earth. This is shown in Table II, which gives the analysis of air at different heights above the surface of the earth. Carbon dioxide being the heaviest, is found only in the lower layers of the atmosphere; and nitrogen being the lightest, is practically the only one found in the upper layers. If air were a compound of these gases, it would be of the same composition throughout.
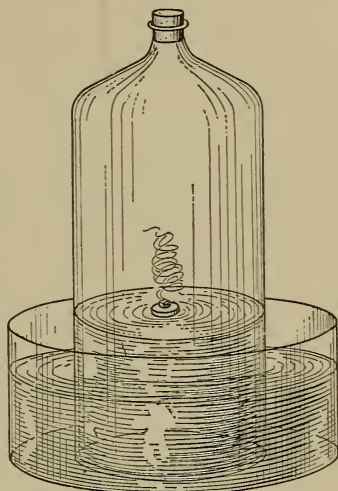
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TABLE II

Composition of the Atmosphere at Different Levels

	Earth's surface per cent	6 miles per cent	12 miles per cent	30 miles per cent	60 miles per cent
Nitrogen.................	78.0	81.1	86.0	89.6	95.3
Oxygen.................	21.0	18.3	13.8	10.3	4.6
Argon.................	0.94	0.58	0.22	0.07	0.00
Carbon dioxide........	0.03	0.02	0.01	0.00	0.00

Carbon Dioxide.—Let us now consider the individual gases of the atmosphere. *Argon* is one of a group of rare gases that are of scientific interest only, and need not be discussed here. *Carbon dioxide* is the product formed when carbon, or any compound of carbon, burns. It is also produced in animal bodies by the combustion of the food materials, and is breathed out of the lungs in considerable quantities. It is commonly called carbonic acid gas. As can be seen in the accompanying table, there is but a very small amount of it in the air, 0.03 per cent. This means that in 10,000 quarts of air there are only 3 quarts of carbon dioxide. The carbon dioxide gets into the air in considerable amount from the burning of fuel. When one considers the millions of tons of coal and other fuel that are burned in the world every year, it may seem remarkable that the amount of carbon dioxide in the air does not increase rapidly. Two things should be kept in mind in this regard, however. The first is, that a few million tons of carbon dioxide really form but a very small fraction of the total bulk of our atmosphere. (See question 5 at end of chapter.) The second is, that plants are continually taking carbon dioxide from the air and using it to build up their tissues; and when it is remembered that the area of the earth's surface covered by green vegetation is thousands of times greater than the area covered by cities and furnaces, it will be readily

understood how nature maintains the balance of carbon dioxide in the air. The amount of carbon abstracted from the air by plants is really enormous. For example, an acre of mature corn contains about 6200 pounds of dry material, and nearly half of this is carbon. Thus 3000 pounds of carbon are abstracted from the air by an acre of this plant in a single season. Two or three heavy cuttings of alfalfa may contain 5000 pounds of carbon per acre. It should be mentioned here that many analyses of air, from many parts of the earth, at various seasons of the year, show but extremely small variations from 0.03 per cent of carbon dioxide. Air over cities usually shows higher percentages than air over the country. The chemistry of carbon dioxide will be discussed more fully in Chapter VI.

Nitrogen.—This gas occupies nearly four-fifths of the air by volume, or three-fourths of it by weight. This is clearly shown in Figure 2. There is thus an enormous mass of it altogether in the air. Since, however, there is only a small trace of it in the solid portion of the earth, that is in the soil, rocks, and water, nitrogen really occupies but 0.02 per cent of the total weight of the earth. Its compounds, however, are of extreme importance; they enter into the composition of every plant and animal cell; egg white, lean meat, the curd of milk are spoken of as nitrogenous foods; practically all high explosives and artificial dyes are compounds of nitrogen; in the soil it is usually found in very small amounts, and is the most costly fertilizer element. All through this course, therefore, we shall have occasion to talk about the compounds of nitrogen.

Properties of Nitrogen.—In the air the nitrogen is not combined as a compound; it exists as free nitrogen. In this condition it is a rather uninteresting element chemically. It is very inert, undergoing chemical changes and

entering into chemical reactions with great difficulty. As was seen in the bell-jar experiment a few pages back, it is a colorless, odorless gas. It will not dissolve readily in chemicals, it will not burn, it will not support combustion, it cannot be experimented with in any way like many other gases which we shall study.

Nitric Acid.—There are a few violent ways in which it can be made to unite with oxygen, the most important of which is the electric spark. When lightning discharges through the clouds in a storm, a considerable amount of nitrogen and oxygen combine to form nitrogen pentoxide, according to this equation:

$$2\,N \ + \ 5\,O \ = \ N_2O_5$$
nitrogen   oxygen   nitrogen pentoxide
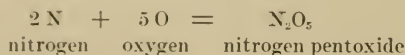
This nitrogen pentoxide then dissolves in the rain water to form nitric acid, thus:

$$N_2O_5 \ + \ H_2O \ = \ 2\,HNO_3$$
nitrogen pentoxide   water   nitric acid
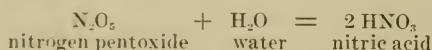
Some ammonia, NH_3, is formed by the electric discharge at the same time, and this combines with the nitric acid to form ammonium nitrate. Both the nitric acid and the ammonia are very valuable and quick-acting fertilizers, and it is these which cause the wonderful greening of grass within a short time after a thunder shower. It has been found that from 5 to 8 pounds of nitrogen per acre are brought down by rain each year in this manner. Engineers and chemists, working together, have devised machines which imitate this process of nature for producing nitric acid. In Norway, where electric power is plentiful and cheap, enormous amounts of nitric acid are made in this way. And it is claimed that Germany could not have continued the World War so many years if she had not utilized this process for making nitric acid, needed in large quantities for making high explosives.

Explosives.—In these nitrogenous explosives advantage is taken of the fact that since nitrogen forms compounds with great difficulty, these compounds are also easily decomposed—by flame or even by concussion. Imagine a barrel of nitrogen gas combined into a solid chemical the size of a baseball; and this chemical enclosed in a cannon barrel and decomposed by a percussion cap. Within a fraction of a second the nitrogen is liberated as a gas, and it tries to expand to its original volume of a barrel; at the same time the great heat developed makes it expand many times more than this. We thus get an idea of the enormous pressure developed behind a cannon ball by the decomposition of an unstable compound of nitrogen.

Free Versus Combined Nitrogen.—Although all plants and animals require compounds of nitrogen as food, they cannot avail themselves of the uncombined nitrogen of the air. We breathe great quantities of nitrogen into our lungs, but it is all breathed out again unchanged and unabsorbed. A potato plant may be starving for nitrogen in a poor soil, and still be unable to utilize the tons of nitrogen blowing around it in the air. One class of plants, the legumes, to which the peas, beans, and clover belong, can make use of atmospheric nitrogen when they have certain bacteria living in little nodules or lumps on their roots. We shall have more to say about this when we come to the question of plant nutrition.

Since nitrogen as an element, then, is so inactive, its chief purpose in the air is to dilute the oxygen. This brings us to the next most abundant gas of the atmosphere, *oxygen*, the "elixir of life," the "vitalizing element."

Combustion.—Everyone knows that any substance must have air in order to burn: exclude air from a fire, and it goes out; exclude air from the lungs of an animal, and combustion in the body ceases. In other words, air supports combustion; it enables a substance to burn when once

afire. Nearly everyone now knows that it is the oxygen part of the air that really supports combustion. This fact was discovered in 1774 by Joseph Priestley, an English chemist, and this discovery of oxygen is one of the great landmarks in the history of chemistry. Combustion is no doubt one of the most important single chemical reactions that is known. As we now know it, *ordinary combustion is union with oxygen, accompanied by the production of heat.* Combustion may vary in speed and violence from the blazing of a splinter or the instantaneous explosion of gasoline vapor, to the slow combustion of the food in our bodies, or the even slower burning of food in the bodies of cold-blooded snakes and fish. Combustion furnishes the world with most of its energy; the combustion of coal, wood, and petroleum products keeps us warm, runs our engines and machinery, hauls us on trains or on boats; and every living animal gets its energy to live and to do work by the combustion of fuel in its body. We cannot appreciate too highly the importance of this universal chemical reaction. We have already given several examples of simple combustion, as the burning of sulfur, of hydrogen, and of phosphorus; more complex examples will be given from time to time.

Oxygen in the Air.—Let us emphasize again here, that it is on the 21 per cent of oxygen in the air that we are dependent for the support of this combustion. If the air were all oxygen instead of only one-fifth, combustion would proceed many times more rapidly. An interesting demonstration of this is the fact that iron and aluminum burn brilliantly in a bottle of oxygen. A mouse put into a bottle of oxygen burns up its fuel faster than it can digest more food for fuel, and it soon dies. Advantage has been taken of this increased respiration in the body with an increased supply of oxygen in the machine called

the pulmotor, which forces oxygen into the lungs of persons who have suffocated by water or by gas.
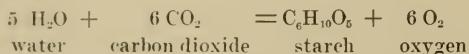
Plants and Carbon Dioxide.—In discussing carbon dioxide, it was pointed out that millions of tons of it are thrown into the air from our furnaces; that it does not accumulate in the air, however, because plants are continually abstracting it to build up their tissues. But why

Fig. 3.—An open fire-place. In this form of heating arrangement only a very small percentage of the developed heat radiates into the room; but it is a cheerful heating plant, and the open flue causes continuous ventilation.

does not the enormous consumption of oxygen in our furnaces soon deplete the supply in the air? Chemical analysis shows that the oxygen is not being depleted; that it remains constant at about 21 per cent of the air. For the answer to this question we must go to the plants, as we did in the case of carbon dioxide. When a plant absorbs carbon dioxide, it is the carbon part of it that is used. The carbon is combined with water to form the

well-known class of substances called carbohydrates (the
name indicating that they are compounds of carbon and
water). Starch, sugar, and cellulose fibres of plants are
examples of such carbohydrates. Let us use starch,
$C_6H_{10}O_5$, to illustrate this chemical reaction in the form
of an equation:

$$5\ H_2O\ +\qquad 6\ CO_2\qquad =C_6H_{10}O_5\ +\quad 6\ O_2$$
$$\text{water}\qquad \text{carbon dioxide}\qquad \text{starch}\qquad \text{oxygen}$$

The equation tells us that six atoms of carbon from the
carbon dioxide combine with five molecules of water to
form one molecule of starch. In taking the carbon out
of the carbon dioxide, there is left the six O_2 groups.
Since the plant has no use for all this oxygen, it exhales
it back into the atmosphere. In other words, for all the
carbon dioxide that the plant absorbs from the air, there
is a corresponding amount of free oxygen returned to it.
In this way the balance between them is maintained.

Photosynthesis.—It should be remembered that only in
the green parts of plants, and only in sunlight, can this
remarkable transformation of carbon dioxide and water
into starch be carried on. *This process is called photo-
synthesis, which means "manufacturing by means of
light."* The plant uses the starch, then, to build up all
the other substances of its body. Hence photosynthesis
is the fundamental chemical reaction behind all crop pro-
duction, forest growth, and coal formation (since coal is
decayed vegetable matter). We can now understand the
tremendous importance of the trace of carbon dioxide
in the air. Since photosynthesis is so important, it will
be mentioned several times in this course. At the present
moment we are interested in it because it is helping to
maintain constant amounts of carbon dioxide and of
oxygen in our atmosphere.

Oxidation.—The chemistry of oxygen is far more inter-
esting than that of nitrogen. Whereas the latter is very

inactive, oxygen is very active. It combines readily with
practically all the other elements. Combined with silicon
in the form of silicon dioxide, or quartz, it constitutes
three-fourths of the weight of the earth. Combined with
hydrogen, it forms water, which makes up three-fifths of
our bodies, and covers two-thirds of the area of the globe.
*The chemical reaction involved when anything combines
with oxygen is called oxidation. Hence combustion is*
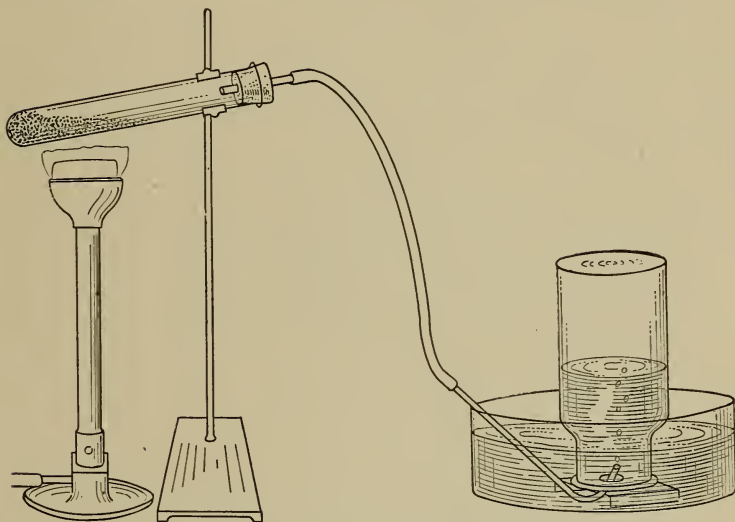
Fig. 4.—Apparatus for the preparation of oxygen.

rapid oxidation. Oxidation, the compounds of oxygen,
fuels, and related subjects, are so very important that a
separate chapter is reserved for them.

As regards the other characteristics and properties of
oxygen, we know that it is odorless, colorless, and taste-
less; Table II shows us that it is a little heavier than
nitrogen, but lighter than carbon dioxide. It dissolves
in water to an appreciable extent, and it is this dissolved
oxygen that fish utilize when they "breathe water."

Other Constituents of the Air.—We have now discussed argon, nitrogen, ammonia, nitric acid, carbon dioxide and oxygen as being the constituents of our atmosphere. Besides these, we know that there is water vapor in varying quantities, and that there is a continual transfer of water from the earth to the air by evaporation, and from the air to the earth as rain. Since, however, rain, snow, clouds, dew, etc., are physical phenomena and not chemical, they cannot be discussed here. Very often there are abnormal constituents of the air due to our civilization : fumes from factories, smokestacks, and smelters are sometimes abundant enough to cause harm to both animals and plants. There are also always dust and bacteria in the air.

QUESTIONS

1. Make a list of all the substances found in the atmosphere, giving the amount of each that is present (whenever given in the text), and stating which are compounds and which are elements.
2. How do we know that air is a mixture of these things and not a compound of them (two proofs) ?
3. Write equations showing the chemical reaction involved when oxygen combines with carbon, hydrogen, sulfur, phosphorus, and nitrogen.
4. Explain in detail why the amounts of carbon dioxide and of oxygen in the air remain the same.
5. By means of the data given in this chapter, compute the amounts of carbon dioxide, in pounds, over an acre of land.
6. What is combustion ?
7. Describe the element nitrogen.
8. Why is the nitrogen of the air of no use to most plants ?
9. What is nitric acid ? What is it used for ? How is it made ?
10. Explain the process of photosynthesis, using a chemical equation.
11. Why is photosynthesis important ?
12. Describe the chemical and physical characteristics of oxygen.
13. Name eight compounds of oxygen mentioned so far in this book, giving both chemical name and common name where possible.

LABORATORY EXPERIMENTS

6. To Show the Presence of Carbon Dioxide in the Air.—Almost fill two watch glasses with clear limewater. Immediately cover one of them with a glass plate, and leave the other exposed. After a half hour or more examine the surface of each for a cloudy precipitate or sediment. If either of them shows the cloudiness, it means that carbon dioxide has had access to it, as carbon dioxide always turns limewater milky. (See page 101 for the equation for this reaction.) Why was one watch glass covered and the other not ? Prove the presence of carbon dioxide in the breath.

7. To Show the Presence of Water Vapor in the Air.—How do the following commonly observed facts show that there is water vapor in the air: the "steaming" of eye-glasses on coming into a warm room, the "sweating" of cold water pipes and water pitchers in summer time? How does the water vapor get into the air? What are clouds? How does rain form?

8. To Prepare Nitrogen From the Air.—Set up the apparatus shown in figure 2 with a piece of phosphorus the size of a pea on the cork floating on the water. (CAUTION: handle phosphorus with forceps!) A glass dish or a pan will do for holding the water. Ignite the phosphorus or candle, then cover with a quart Mason jar, or with a bell jar, tightly stoppering the latter after it is in place. Allow to stand until the burning ceases, and all the phosphorus fumes have settled and dissolved in the water. Observe the difference in water level inside and outside the jar. What caused this? What is the principal gas left in the jar? Thrust a burning splinter into it. Explain the result. If all the oxygen had been removed, how much would the air inside of the jar have shrunk? What was formed by the burning phosphorus?

9. The Preparation and Study of Oxygen.—(Two students working together.) Set up the apparatus shown in figure 4. Place in the test tube a mixture of 5 g. potassium chlorate, $KClO_3$, and 1 g. of manganese dioxide, MnO_2, which have been ground together in a mortar. Heat slowly along the whole length of the tube, holding the burner in the hand. When the materials melt and begin to boil, it means that oxygen is being given off. At this point regulate the heating so as to just keep the reaction going. Pass the gas into the inverted bottles of water. When a bottle is filled, cover its mouth with a glass plate, lift from the water, and begin the filling of another bottle. When the reaction is over, make the following tests on the bottles of gas:

(a) Thrust a glowing splinter into one of the bottles. How does pure oxygen affect the rate of combustion? How would it affect an animal breathing it? What is a pulmotor?

(b) Thrust a glowing piece of charcoal held by tongs into one of the bottles and cover it with a glass plate. When the burning has ceased remove the charcoal, put a little water into the bottle and shake vigorously. Pour the water into a little clear limewater in a test tube. What was formed by the union of the carbon with the oxygen?

(c) Fill a deflagrating spoon with sulfur, ignite and lower it into a bottle of oxygen. When it ceases to burn, remove the spoon, shake the contents of the bottle with a little water and then test the water with litmus paper. (See page 142 for the reactions involved.)

(d) Heat a piece of fine bright iron wire red hot and quickly thrust it into a bottle of oxygen. What product is formed here?

Write equations showing how carbon, sulfur, iron and magnesium combine with oxygen. What are oxides? How are they formed? The following equation represents the chemical reaction in the preparation of oxygen:

$$KClO_3 \quad = \quad KCl \quad + \quad 3\,O$$

<div align="center">

potassium   potassium   oxygen
chlorate   chloride

</div>

The manganese dioxide is added to speed up the reaction; it does not take part in it directly.

CHAPTER IV

WATER

Importance.—We saw in the last chapter that one of the commonest and most important of the many compounds of oxygen is water. Until we stop to reflect, we do not realize how profoundly water enters into our daily lives. Three-fifths of our body is water; from 75 per cent to 90 per cent of all our fresh meats and vegetables is water; milk has 87 per cent of water; lettuce and cucumbers as high as 93 per cent; the driest hay is one-tenth water; people in the country use from six to eight gallons of water per person per day for ordinary household purposes, and in the city, with street sprinkling, sewerage and factories, there is four or five times this amount used. The rainfall over most of our average humid states amounts to from 30 to 50 inches a year. It takes from 400 to 800 pounds of water to produce each pound of dry matter in our crops. There is more water surface than land on our globe; and in many places the ocean is from five to six miles deep. We know that all forms of animal and plant life are dependent on water for their continued existence. One of the surest ways of killing a plant or animal is to keep water away from it. And one of the most effective means of preserving meats, fruits and vegetables is by drying, since microörganisms cannot ferment or spoil them with so little water present. It should be pointed out here that the great majority of chemical reactions take place in the presence of water. This might be inferred from the statements above, for all life processes are series of chemical reactions. In many chemical reactions water is not only present, but it also

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actually enters into the reaction. The conversion of starch to sugar by the saliva of the mouth is such a reaction; thus:

$$C_6H_{10}O_5 \quad + \quad H_2O \quad = \quad C_6H_{12}O_6$$
$$\text{starch} \qquad \text{water} \qquad \text{sugar}$$
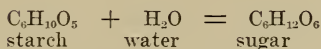
This kind of reaction will be taken up more fully under the chemistry of digestion. Since water is so universally important, it is desirable that we take up at this time the chemistry of water.

Composition of Water.—As has been said, water is a compound of two elements, hydrogen and oxygen. This can be proved in two very interesting ways. The first is to decompose water, and prove that the products formed are hydrogen and oxygen. The second is to take hydrogen and oxygen, cause them to combine, and then note that the combination formed is water.

Electrolysis of Water.—The decomposition of water can be brought about in several ways, but the most effective and easiest is by means of an electric current. It is done in the apparatus shown in figure 5. It is essentially two tubes of water, closed at the top, and connected at the bottom; the current passes into one tube, through the water to the other tube, then out to the source of the current. The water is acidified with a little sulfuric acid in order that it may conduct the current more easily. The current breaks up the water into the two gases, and, due to certain laws of electro-chemistry which we cannot go into here, each gas rises in a different tube and collects at the top. After considerable gas has been generated, it will be noticed that there is twice as much in one tube as in the other. Which gas is oxygen and which is hydrogen? From the experiments on oxygen, the student will remember that pure oxygen causes such rapid combustion that a small glowing coal on the end of a splinter will burst into flame when thrust into the oxygen. This test
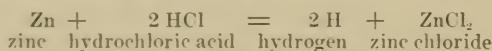
can now be applied to the gases in the tubes, as the tubes
are supplied with cocks at the top for letting out the gas.
It will be found that the smaller volume of gas is the
oxygen. The hydrogen can best be tested by making
use of its inflammability, as hydrogen burns very readily.

Thus we have proved that water can be decomposed
into two elements, hydrogen and oxygen. Furthermore,
the above experiment shows that there are two volumes of
hydrogen to one of oxygen. The equation for this reaction
would be as follows:

$$\underset{\text{water}}{H_2O} \;=\; \underset{\text{hydrogen}}{2\,H} \;+\; \underset{\text{oxygen}}{O}$$
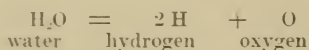
It will be noticed that neither the electric current nor the
acid enters into the equation; they are simply part of the
machinery that does the work.

Hydrogen.—Before taking up the second proof of the
composition of water, that is, by combining hydrogen and
oxygen, we should take a moment to discuss this new
chemical element that has come to our attention, namely,
hydrogen. For purposes of study and experiment it can
be prepared in several ways. One way, of course, is from
water by the electric current. A still easier way is by
allowing some acid, as hydrochloric, to react with some
metal, as zinc, in an apparatus like the one in figure 11.
As its name indicates, hydrochloric acid is a compound of
hydrogen and chlorine; and when the acid reacts on zinc,
the zinc displaces the hydrogen. Thus:

$$\underset{\text{zinc}}{Zn} + \underset{\text{hydrochloric acid}}{2\,HCl} \;=\; \underset{\text{hydrogen}}{2\,H} + \underset{\text{zinc chloride}}{ZnCl_2}$$

The hydrogen is generated in a steady stream from the
nozzle, and can be collected in bottles or burned at the
nozzle. Again, if steam is passed over red-hot iron shav-
ings in a tube, the iron combines with the oxygen of the
steam and liberates the hydrogen.

Properties of Hydrogen.—However it is prepared, hydrogen is found to be a colorless, odorless gas, like nitrogen and oxygen. Unlike nitrogen, however, it is very active chemically; and unlike oxygen it does not support combustion, but burns very readily. In fact, burning hydrogen forms one of the hottest flames known. In burning, of course, it unites with oxygen; and the so-called oxy-hydrogen blow-pipe is a device for producing a small pencil of flame for heating small areas of metal for welding and cutting; the limelight in stereopticons and theatres is produced by a piece of lime heated white hot by an oxy-hydrogen flame.

Hydrogen is the lightest gas known, being only one-fifteenth as heavy as air. It therefore finds great use in the filling of balloons and airships. This property can be demonstrated in the laboratory by blowing soap-bubbles with the stream of hydrogen from the generator. Because of its lightness, there can be no hydrogen in our atmosphere, unless it be in the very top layers, many miles above the earth.

FIG. 5.—Apparatus for the electrolysis of water. An electric current is passed through the water contained in the tubes. The water is decomposed into hydrogen gas, which occupies the negative tube, and oxygen gas, which occupies the positive tube.

Compounds of Hydrogen.—There are many important compounds of hydrogen. For example, all acids contain hydrogen; all petroleum products, fats and oils, the carbohydrates, ammonia, acetylene, alcohol, are compounds of hydrogen. They will be met with further on in their appropriate places in the book.
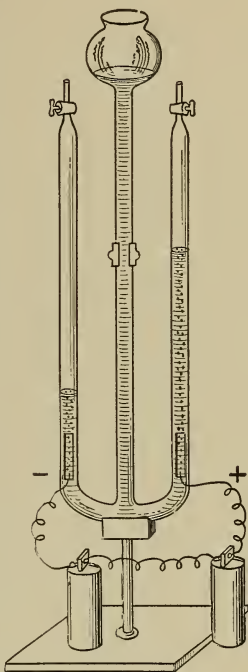
Synthesis of Water.—Going back now to the proposition that hydrogen and oxygen unite to form water, how can we prove that they do? A very simple proof is as follows, based on the well-known fact that water vapor condenses to a liquid on cold surfaces: if a test tube or beaker full of cold water, but perfectly dry on the outside, is held a half inch or so above a flame of burning hydrogen, in a moment the sides of the tube will be covered with a mist of fine water drops. This water originated in the jet of burning hydrogen, and on striking the cold surface condensed to a liquid. The same proof is often furnished when a tub of cold water is put over a gas or kerosene flame; the bottom and sides of the tub "sweat" profusely until the tub becomes too warm to condense the water vapor. These last cases also show that hydrogen is present in considerable quantities in both coal gas and kerosene; in the former it is present as the free element, and in the latter in a compound.

From the above experiments and observations we can draw two conclusions: First, that when hydrogen gas burns, it forms water; second, that when a compound of hydrogen burns, the hydrogen part of the compound forms water. The chemical equation involved is as follows:

$$2\,H \quad + \quad O \quad = \quad H_2O$$
hydrogen   oxygen   hydrogen oxide (water)
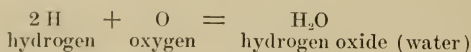
Hydrogen Peroxide.—It has been said above that when hydrogen and oxygen combine, it is in the ratio of 2 to 1 by volume. This is true in ordinary burning of hydrogen. Under special conditions, however, these two elements can be made to unite in equal parts, that is, two volumes of each:

$$2\,H \quad + \quad 2\,O \quad = \quad H_2O_2$$
hydrogen   oxygen   hydrogen peroxide
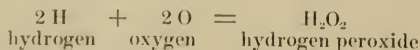
This hydrogen peroxide is the well-known disinfectant known also as "peroxide" and "dioxygen." It is also

a bleaching agent. It owes both its bleaching and germi-
cidal effects to the fact that the extra oxygen atom is easily
broken off, and this atom of oxygen oxidizes away colors
and kills germs. Only water is left behind; thus:

$$H_2O_2 \quad = \quad H_2O \quad + \quad O$$
$$\text{hydrogen peroxide} \quad \text{hydrogen oxide} \quad \text{oxygen}$$
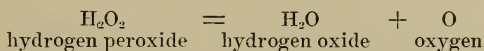
Water as a Solvent.—If water is simply the oxide of
hydrogen, the question now arises, how can we have good
water and bad water? By good and bad one usually means
good or bad for a certain purpose; a water may be good
for drinking, but bad for laundering. We also have soft
and hard waters, mineral and carbonated waters, and
others. As a matter of fact, water itself is the same the
world over, just the same as any chemical compound
always has exactly the same composition. But water has
a remarkable capacity for dissolving and absorbing other
substances. We know that we can dissolve large quanti-
ties of salt in water to form a brine; that sugar will readily
dissolve in water to form syrup; that most medicines are
taken in water; that water dissolves soap, also the dirt and
stains on clothing. Many minerals dissolve in water; in
fact, plants get all their mineral food from the soil by
means of the soil water passing up through the roots and
stems. In short, water has been called "the universal
solvent," since practically everything will go into solution
in water in some degree. Glass, sand, silver, porcelain,
will all dissolve in amounts sufficient to be detected by
chemical means.

Impurities in Water.—No wonder, then, that water is so
often impure; it is practically never free from some im-
purities. As rain falls through the air, it dissolves the
nitric acid and ammonia formed by the lightning dis-
charges; it takes up sulfur fumes, thrown out from smoke
stacks; dust and bacteria are brought along to earth, so

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that even rain-water contains various foreign substances. Rain over the sea, or rain falling after an hour of heavy raining, is very free from impurities. As the water seeps through the soil it absorbs a great number of minerals, organic matter, particles of dirt, bacteria, gases, etc.; hence the usual appearance and taste of our river waters. If the water seeps through layers of sand and rock, most of the solid particles and bacteria are removed from it; hence well water and spring water are usually clear. But this seepage through rocks does not remove mineral matter that is in solution; hence well and spring waters may, and usually do, contain considerable mineral matter, especially lime and magnesium compounds. In this same way water that falls onto soil contaminated with manure, decaying animal and vegetable matter, etc., becomes contaminated; and many times wells are situated down hill from such places, and the contaminated water enters the wells and spreads disease germs to everyone using it. Such water often has a bad smell and taste, and is yellowish in color because of the organic material in it.

Hard Water.—The commonest impurities in water are compounds of calcium and of magnesium, principally the carbonates and sulfates. Water containing these substances is called "hard water." These salts get into the water when it flows through limestone formations. It is perfectly good for drinking, but when used for washing purposes the lime and magnesia combine with the soap and form a sticky curd, that is very difficult to remove from either clothes or tubs. There are several ways of removing these minerals or "softening" the water; but as the chemistry involved would be understood with difficulty at this place, it will be taken up in Chapter VIII.

Purification of Water.—The various impurities in water can be classified conveniently, according to the methods of removing them, as follows:

1. Solid particles in suspension.
2. Solids in solution.
3. Gases.
4. Bacteria.

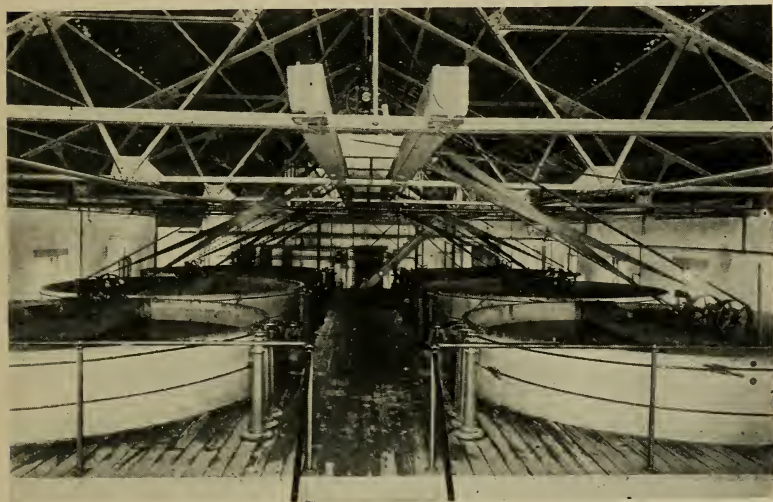A discussion of each of these kinds of substances and

Fig. 6.—View in a water filtration plant that has a capacity of 3,000,000 gallons a day. The water filters by gravity through beds of fine sand 35 to 40 inches deep. This removes all suspended matter, including any sediment produced by the chemical treatment of the water.

the ways of removing them from water will now be taken up.

1. *By solid particles in suspension* is meant such things as fine sand and clay, pieces of plant material, and other substances. They make the water cloudy, and the particles can often be seen by the naked eye. These substances are removed by a process called *filtration*. In the laboratory this is usually done by folding a special kind of paper, similar to blotting paper, so that it will fit the inside of a funnel, and then pouring the muddy water into the paper. The water can pass through the fine pores of

the paper, but the particles of dirt cannot; they are retained in the paper and the clear water flows out of the funnel. In the purification of river and lake waters for use in cities, this same principle is utilized on an enormous scale (Fig. 6). Instead of filter paper, beds of fine sand from 25 to 40 inches deep and covering many square yards
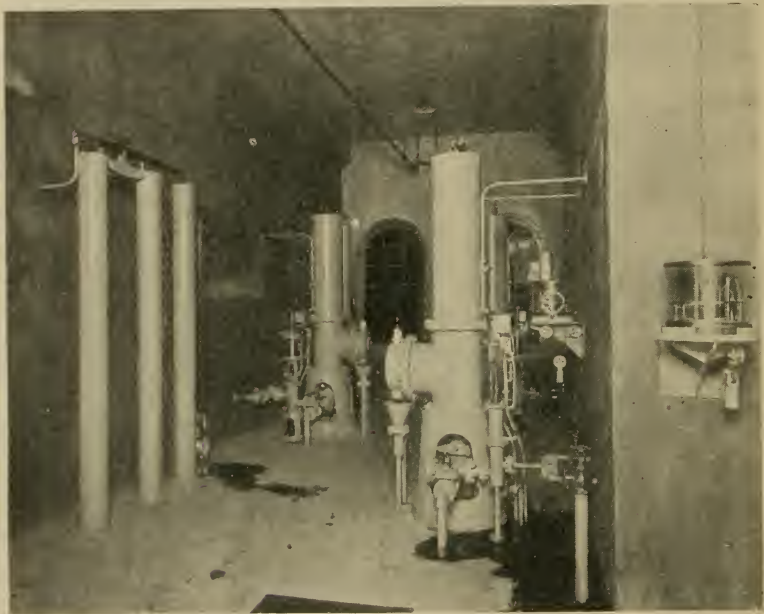
Fig. 7.—Automatic machines for feeding chemicals into the water in a water purification plant. Lime is added for softening, alum to coagulate the organic matter, and hypochlorite of lime to kill bacteria.

or even acres are used as the filtering medium. Tile drains at the bottom of the beds carry off the clear water. In most cases the water is treated chemically with alum, in order to coagulate the coloring matter and other organic impurities, and these are then left on the sand, together with the dirt (Fig. 7). Water softeners are sometimes added at the same time. In this way very good water for

millions of people is obtained even from such rivers as
the Mississippi, the Seine, and the Thames.   Brick filters
are often installed in household rain-water cisterns to
remove dirt.

2. *When solids are in solution* in water the particles
of the solid cannot be seen; the water is perfectly clear.
Examples of this class would be sugar, salt, lye, the lime
in hard water, the alkali in "alkali water."   It will be
readily agreed that filtration would not remove any of
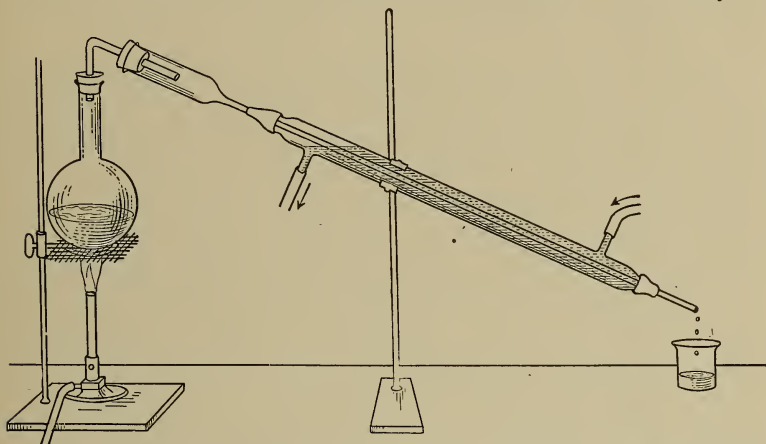
Fig. 8.—Apparatus for the distillation of water.   Water is boiled in the flask at the left,
and the vapors pass into the long water-jacketed tube, where they condense to a liquid and
flow into the receiver at the right.

such substances.   Neither can many of them be removed
by chemical means.   Hence, another process is employed,
called *distillation*.   A simple glass laboratory distillation
apparatus is shown in figure 8.   The liquid to be distilled,
water, for example, is placed in the flask at the left and
heated to boiling.   The steam can escape only through
the side tube which leads into the condenser.   This is a
double-walled tube, the space between the walls being
filled with cold running water.   As the steam enters the
inner tube it comes in contact with the cold walls, and

is condensed into water, which flows out into the receiver at the right.   *Distillation, then, is simply the boiling of a liquid and the condensing of the vapor back into a liquid.* Thus anything which can be vaporized can be distilled.

It can be readily seen that if water containing solids in solution, such as salt, sugar, or alkali, be put into such an apparatus and distilled, the solids will be left behind in the distilling flask; the distilled water will be entirely free from them.   Solid particles in suspension, as discussed in paragraph 1, p. 35, would, of course, also be left behind.

Importance of Distillation.—Distillation is of very great industrial importance.   By this process alcohol is removed from fermented liquor, and we thus get our distilled beverages and our pure alcohol.   Perfumes are removed from flowers by distillation; the nicotine is distilled from waste tobacco and used as an insect poison; turpentine, many medicinal oils, carbolic acid, tar, ether, chloroform, are all obtained by this method.   But no doubt the largest distillation industry in the world is that of petroleum. When the crude oil is pumped from the ground, it is a mixture of dozens of different substances, each of which vaporizes at a different temperature.   When the oil is placed in a distilling apparatus and heated, the substances of lowest boiling point vaporize and condense first, those of highest boiling point last.   In this way successive fractions are distilled over, and we have gasolene, naphtha, benzene, kerosene, paraffine, lubricating oils, and many others.   Millions of barrels of oil are distilled yearly in this country.   Long rows of stills, each holding as high as 1200 barrels, are operated at one time.   They are fed from the oil fields by pipe lines that in many cases are over 1000 miles long.

3. *Many gases* dissolve in water; air to a considerable extent, carbon dioxide to a still greater extent, and the

gases formed in decomposing vegetable and animal matter in sewerage, soil, manure, etc., to an extent sufficient to make them very disagreeable to smell and taste in many cases. Their removal, however, is very simple, and can be illustrated by a simple experiment. Some odorous gas, such as hydrogen sulfide or ammonia, is dissolved in a beaker of water and the latter boiled. A few minutes of boiling will entirely remove the odor of the dissolved gas.
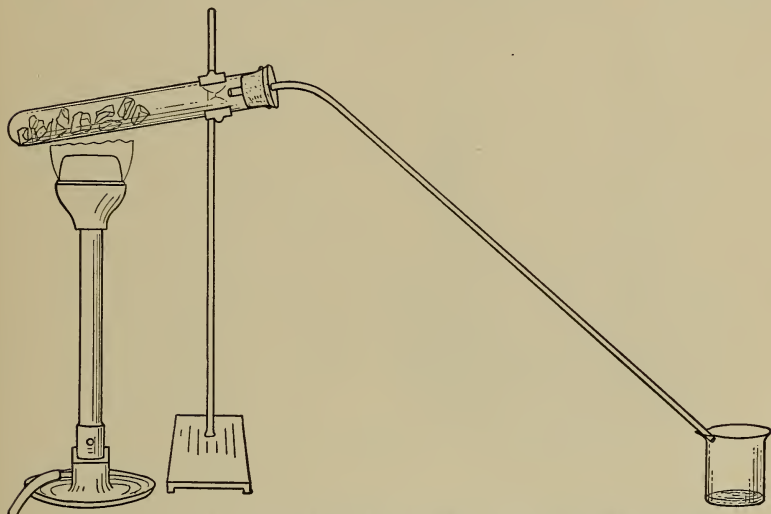
Fig. 9.—Apparatus for demonstrating the presence of water in fresh plant tissue.

This is because *the gas is not soluble in boiling water.* When a tub of water begins to warm up on a stove, bubbles of gas appear on the bottom and sides long before the water begins to boil; these bubbles are the air that was dissolved in the water. This can be demonstrated by placing a minnow in some boiled and cooled water; it will soon suffocate because of lack of oxygen. (See experiment 15.) Thus, gaseous impurities can be removed by boiling. Although distillation also involves boiling, the gases will redissolve in the cooled condensed vapors,

and the distillation will not succeed in removing them.

4. *Bacteria* are minute one-celled plants that are found practically everywhere in countless myriads. Many of them are harmless; many are very beneficial; but some of them are the causes of our most serious diseases. A few disease-producing organisms occur in polluted water, and in this way find their way into the bodies of man and animals. Typhoid, cholera, and dysentery are among

our most dreaded water-borne diseases. Hence, it is very often important that we have means of killing or getting rid of them. Fortunately, heat is our most effective disinfectant, and can be used on a small scale in the home for both water and milk. It would be far too costly, however, for purifying the water for a city.
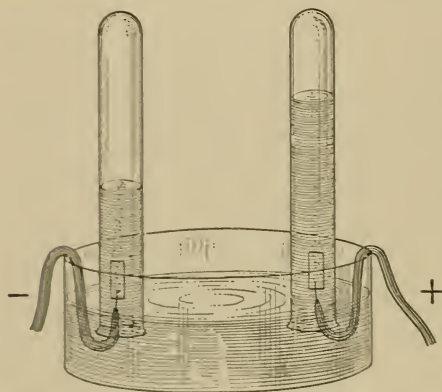
Fig. 10.—Home-made apparatus for the electrolysis of water.

Since a bacterium is a solid plant body, although a very small one, it belongs to the class of substances discussed in paragraph 1, and can be removed by filtration, provided the pores of the filter bed are fine enough to retain them. This method is universally used in the big filtration plants. Chemicals, of course, such as corrosive sublimate, carbolic acid, hydrogen peroxide, formaldehyde, and alcohol, are commonly used for killing bacteria in wounds, on implements and clothing, in rooms, etc., but obviously cannot be put into drinking water for that purpose. There are, however, two very effective germicides, which can be used with perfect safety in city water supplies, because

the active agent is completely used up in the process. These two reagents are hypochlorite of lime and chlorine. The element chlorine is really the active constituent in both of them. It requires only a few pounds of either of these substances to sterilize a million gallons of water. The chemical reaction involved is essentially as follows:

$$2\,Cl \quad + \quad H_2O \quad = \quad 2\,HCl \quad + \quad O$$
chlorine   water   hydrochloric acid   oxygen
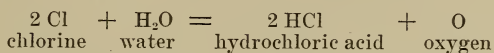
Thus the chlorine does not kill the bacteria directly; it
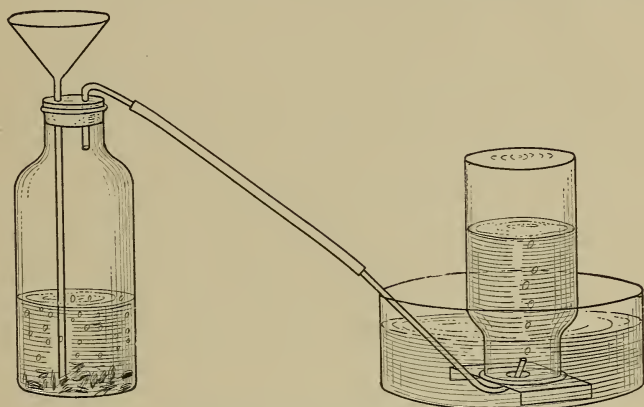
Fig. 11.—Apparatus for the preparation of hydrogen.

liberates an atom of oxygen from the water and this does the work. Hypochlorite of lime is ordinary bleaching powder; in this case its action is the same, the oxygen formed serving to oxidize away the color. In swimming pools, where not much water is supposed to be taken into the body, copper sulfate, blue vitriol, is often used as a disinfectant.

The above cases cover all the ordinary kinds of impurities that are dealt with in water. There are many special industrial uses for water where other problems of purification are met with; for example, in laundries, breweries, chemical works, and steam boilers.

QUESTIONS

1. Describe two different methods for proving that water is composed of two elements.
2. Compute approximately how much water you take into your body in food and drink every year.
3. Describe hydrogen, giving both its physical and chemical properties.
4. What two great uses has hydrogen gas?
5. Tell why rain water is really distilled water. What impurities may rain water carry?
6. Describe the course that water follows before it arrives in our well for drinking, and state the impurities it may have acquired.
7. What is hard water?
8. How could you prepare good water for laundering from a mud puddle? How prepare good drinking water from it?
9. What is filtration?
10. What is distillation? Name ten substances that are manufactured by a process of distillation.
11. What are bacteria? How can they be harmful? What three general methods are there for removing them from water or killing them?
12. How can you prove that air dissolves in water?

LABORATORY EXPERIMENTS

10. **To Prove the Presence of Water in Fresh Plant Tissues.**—Set up the apparatus shown in figure 9, filling the test tube with lettuce or cabbage leaves, or pieces of potato or apple. Heat the tube very gently with a low flame, avoiding scorching of the material. What liquid forms in the long tube? Taste it, and let some of it drip into some blue vitriol that has been heated until it is white (experiment 5). What does this prove? The water has been removed from the plant tissue by distillation (p. 37). What is distilled water?

11. **The Electrolysis of Water.**—(To be done by instructor.) Fill the apparatus with 5 per cent sulfuric acid (since water alone does not conduct electricity sufficiently well). Be sure that the tubes are completely filled. Any source of direct current will do which gives 4 to 20 volts. Four dry batteries work very well, although slowly. Note the relative volume of gas formed in the tubes. Allow the gas of smaller volume to escape through the stop-cock onto a glowing splinter. Explain the result. Ignite the gas from the other tube with a match. Does this test apply to hydrogen only, or would some other gases burn? What does this experiment prove concerning the composition of water?

If regular apparatus for the electrolysis of water is not obtainable, the apparatus shown in figure 10 can be set up. The electrodes can be of copper, although platinum is preferable.

12. **The Preparation and Study of Hydrogen.**—(To be done by instructor.) Set up the apparatus shown in figure 11. Place about 10 g. of flake zinc or iron nails in the flask, place the stopper in firmly, and pour dilute hydrochloric or sulfuric acid down the funnel. Allow the gas to generate and to escape into the air for at least *five minutes by the clock*, so as to force all air out of the flask, replenishing the acid if the generation of gas slows up. Then pin a stout towel around the generator. *These precautions must be observed in order to prevent explosions.* The hydrogen

may now be collected in wide-mouth (never in narrow-mouth; why?) bottles. Make tests on the gas as follows:

(a) Test the inflammability of the gas in one of the bottles.

(b) Light the hydrogen as it issues from the generator. Hold a test tube or beaker filled with cold water just above the hydrogen flame for a few seconds. What forms on the outside? What is its source? Write an equation showing the reaction when hydrogen burns. What is the oxy-hydrogen torch and for what is it used?

(c) Holding a bottle of hydrogen upside down, thrust a burning splinter into the bottle clear to the bottom. Does hydrogen support combustion? Would a jet of oxygen burn in a room full of hydrogen, the same as a jet of hydrogen burns in a room full of oxygen? Why?

(d) Test the lightness of hydrogen by holding the tube from the generator in a bottle right side up, and a bottle upside down, and then testing each bottle for hydrogen with a match. Is hydrogen lighter or heavier than air? What practical use is made of this fact?

13. To Show Various Degrees of Solubility in Water.—Into separate test tubes containing 5 c.c. of distilled water add 5 g. of $CaSO_4$, $CaCO_3$, and $CaCl_2$, respectively. With the thumb over the test tube shake each of them thoroughly for at least a minute. Filter each solution. The filtrates are saturated solutions of the various salts. Place five drops of each on watch crystals and evaporate to dryness. Compare the amount of residue on each.

14. To Demonstrate the Solids Dissolved in Various Waters.—Fill separate watch glasses with (a) distilled water, (b) rain water, (c) hard water from a well or city main, (d) river or other muddy water, and (e) the same water as in (d) but filtered through double filter paper. Evaporate to dryness, and compare the amounts of residue. What is the source of the residue in (b)? Put a few drops of dilute hydrochloric acid on the residue in (c), and also on a little powdered limestone. If the residue in (c) effervesces, what does this signify as to its composition? What is the difference between the kinds of substances in the residues of (d) and (e)?

15. To Show the Presence of Dissolved Air in Water.—(a) Warm some tap-water in a beaker. What are the bubbles of gas which form on the sides and bottom? How do you know they are not bubbles of steam?

(b) If two minnows or gold fish are obtainable, test the solvent power of water for air as follows: Boil a quart of water for several minutes, then completely fill a large beaker with it while still hot. Cover with a glass plate and allow to cool. Meanwhile fill another beaker with water that has been aërated by being poured violently from one dish to another. Cover this also with a glass plate. Now place a minnow in each beaker and watch their behavior for some time. Which first shows signs of an insufficient supply of air? What did the boiling do to the water?

16. To Show the Process of Distillation.—(To be done by instructor.) Distill some muddy water, and then some vinegar. Note the appearance, taste, and odor of the distillate. From what sorts of impurities can water be freed by distillation? What kinds of substances can be distilled?

CHAPTER V

COMBUSTION AND FUELS

WE have now studied briefly the various gases found in the atmosphere; we have seen that the nitrogen is of little chemical importance, but that compounds of nitrogen are of extreme importance in plant and animal life and in certain industries; that the carbon dioxide of the air, although present in very small amounts, is the sole source of carbon for all plants, and that tons of carbon are taken from the air by each acre of crops; we have seen that oxygen occupies one-fifth of the air, and supports all combustion on the earth. We have only briefly discussed the immense magnitude and importance of this process of combustion; and since it is a far bigger subject than our discussion would indicate, we will take it up more in detail now.

Combustion, as we said, is rapid union with oxygen, accompanied by the production of heat. Therefore, *combustion can also be called oxidation*. There are a few exceptions to this where combustion does not involve oxygen, but they need not be considered here. Practically all of our ordinary combustion is the burning of fuel in one of two places, either in engines to give energy for running machinery, or in animal bodies to keep them warm and to give them energy to do work. Therefore the chemical union of oxygen with fuel supplies all the heat and mechanical energy of the world, except the little that we receive from wind, falling water, and the sun. That the sun is really the original source of our fuel will be seen in a later chapter.

A study of combustion, then, is a study of oxygen and

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of fuel. We have already considered the chemical charac-
teristics of oxygen, and we have mentioned that carbon is
the most important fuel element. They unite to form
carbon dioxide:

$$\underset{\text{carbon}}{\text{C}} \quad + \quad \underset{\text{oxygen}}{2\,\text{O}} \quad = \quad \underset{\text{carbon dioxide}}{\text{CO}_2}$$
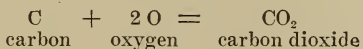
Besides carbon, hydrogen is also a very important fuel
element, either free or combined with carbon:

$$\underset{\text{hydrogen}}{2\,\text{H}} \quad + \quad \underset{\text{oxygen}}{\text{O}} \quad = \underset{\text{water}}{\text{H}_2\text{O}}$$
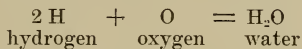
Let us keep in mind from the beginning that carbon diox-
ide and water are always formed when carbon and
hydrogen burn, whether these elements exist free or in
compounds. An interesting fact that should be noted
here is that whereas both carbon and hydrogen as *elements*
may be burned in furnaces and engines, they can be
burned only as *compounds* in animal bodies.

Kindling Temperature.—It is a commonly observed fact
that before a substance will burn it must be heated to a
certain temperature. We find it impossible to fire coal
with a match; hence we use the match to light a piece of
paper, the paper lights some wood, and the wood starts
the coal to burning. *The temperature at which the com-
bustion of a substance will commence is called the kindling
temperature.* The kindling temperature of the paper is
lower than that of the wood, and this lower than that of
the coal. Iron does not ordinarily burn; but if it is heated
very hot it oxidizes on the surface and sluffs off. The
kindling temperatures of gases are usually lower than
those of solids.

Flames.—When a gas burns, a flame is always pro-
duced; *hence we can define a flame as being the combustion
of a gas.* Also we can say that wherever there is a flame
there is a burning gas. This may seem contradictory at

first, since we have all seen both liquids and solids burn
with flames. The explanation of this, however, is that
when kerosene burns, for example, it is the *vapor* of the
kerosene that is burning and producing the flame. The
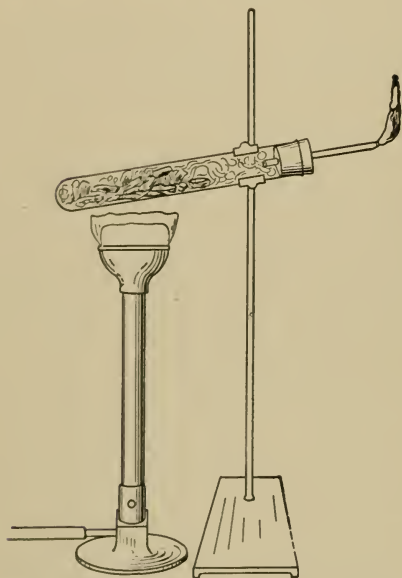same is true of alcohol, gasolene, or other liquid fuel; they
are first vaporized by
the temperature of the
air or the match which
lights them, and then the
vapor burns. When a
kerosene stove is "gen-
erated," it means that
some of the iron parts
must first be heated by a
little kerosene burning
in a special cup, so that
when the regular burner
is lighted, the heated
iron will vaporize the
kerosene as it flows into
the burner, a n d thus
produce a vapor for
combustion.

Fig. 12.—Apparatus for demonstrating the de-
structive distillation of wood. Pieces of hard wood
are put into the test-tube and then roasted. The
gaseous products that are driven off from the wood
escape from the vent, where they can be burned.
After no further gas is evolved, charcoal remains
in the tube. Commercially, the gaseous substances
are not burned, but are made to yield wood alcohol,
acetone, and acetic acid.

The Burning of Wood
and of Coal.—In the case of
a piece of wood burning
with a flame, the origin
of the gas is not quite so
clear. It can be shown,
however, as follows: a few pieces of wood are put into a
test tube, the mouth of the test tube is fitted with a cork
containing a small glass tube, and the wood in the test tube
is gradually heated with a burner (Fig. 12). Soon the
wood begins to blacken, then to give off vapors and fumes.
If a lighted match is applied to the gases issuing from the

mouth of the test tube, they will burn readily, and will continue to do so for some time. When they cease burning, and the contents of the tube are examined, they will be found to consist of pieces of charcoal. If these are placed on an asbestos or iron plate, and fired by the burner, they will burn with a glow just like charcoal, and with practically no flame. We can now see that the burning of a splinter of wood proceeds somewhat as follows: The flame of the match heats up the end of the stick until some gases are driven off; these catch on fire and burn readily now without the match. The heat from the burning gas liberates more and more gas from the wood, and we have a flame as long as any gaseous substances remain in the piece of wood. Then the flame dies out and the piece of wood burns simply with a glow. It is now the charcoal, or carbon, part of the wood that is burning; and when this is consumed, the combustion ceases, and nothing but the ash remains. The same process takes place when coal burns; first the gas burns with a flame, then the carbon part burns without flame. And the big difference between soft and hard coal is that the former has considerable gaseous material liberated when it is fired, and burns with a flame, while hard coal is practically all carbon and burns with very little flame.

The *gaseous* and *solid constituents* of wood and coal are of great industrial value. Charcoal has been produced for centuries. Wood is stacked in a loose pile and covered with earth and sod, leaving a few vents for air around the bottom and one in the middle at the top. The amount of air is so controlled that the gases in the wood can be liberated and burned, but that the charcoal will be left behind unburned. The modern practice is to roast the wood in huge retorts, and instead of burning the escaping gases to separate from them many valuable products, such as wood alcohol, acetone, and acetic acid (Fig. 13). The
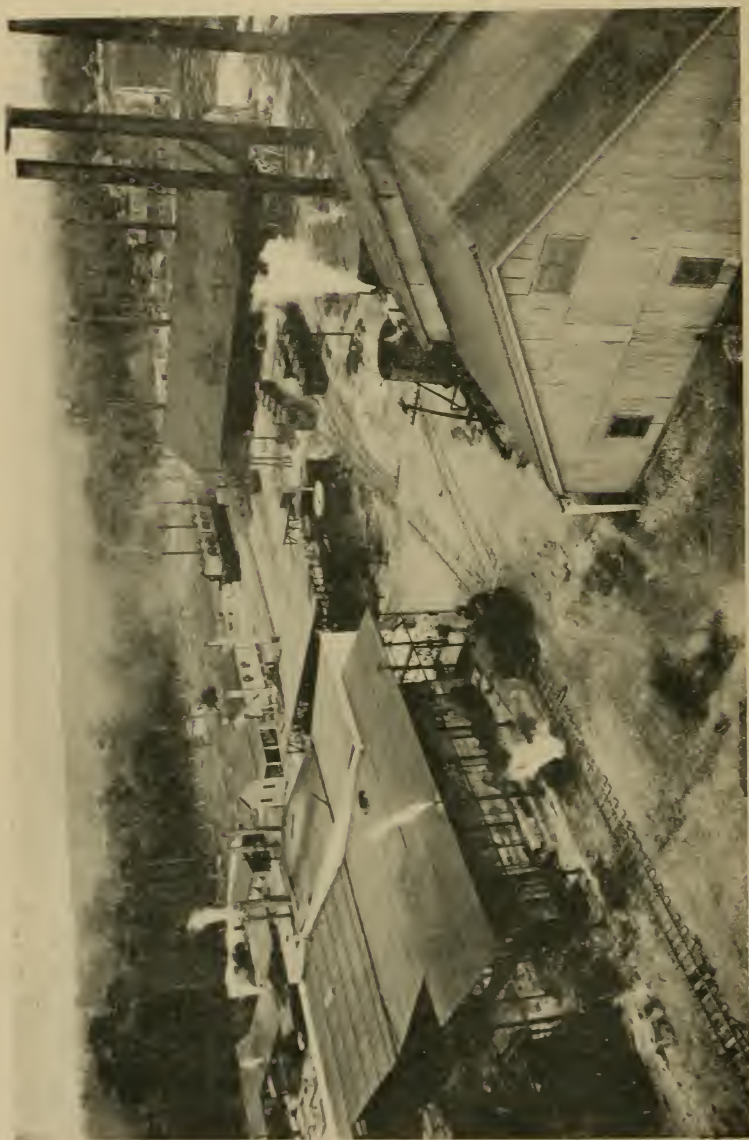
FIG. 13.—Bird's-eye view of a plant devoted to the destructive distillation of wood. Hard woods are the chief ones used, yielding wood alcohol, acetic acid, acetone, creosote, and charcoal. Soft woods yield turpentine, rosin, creosote, charcoal, and tar.

uses of charcoal are discussed in the next chapter. The above process is a certain kind of distillation called *destructive distillation*. It is called destructive, because the composition of the wood is entirely altered; profound chemical changes take place, and substances are formed which never existed in the original wood.

Coal Gas.—The *destructive distillation* of coal is one of our greatest chemical industries. The gas produced, com-
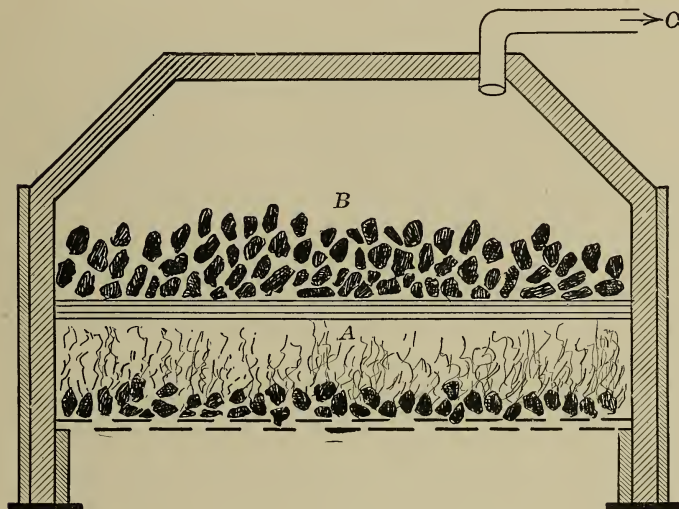
Fig. 14.—Diagrammatic cross-section of a coal-gas oven. Soft coal is charged into the chamber *B*, and roasted by the hot fire in *A*. The gaseous matter driven from the coal passes out at *C* into a long series of machines which remove the tar, the sulfur compounds, the ammonia, and other impurities. The purified gas is then pumped into the city mains.

monly called illuminating or coal gas, supplies most of our cities with fuel for domestic purposes. The coal (always soft coal, because it contains the most gaseous material) is stoked into huge ovens and roasted, air being kept away from it (Fig. 14). The gas which is driven off contains many valuable impurities, such as tar, ammonia, and benzol, and they are extracted before the gas is turned into the

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city mains. From the tar are extracted many compounds important agriculturally, such as cresols for stock dips, creosote for preserving fence posts, carbolic acid, naphthalene for moth-balls, tar and pitch for roofing and road materials, toluol (indirectly used in explosives), and ammonia for fertilizers.

The solid material left behind in the ovens is called *coke*. It consists almost entirely of carbon, together with the ash of the coal. It is used for burning in furnaces to some extent, but its greatest use is in the smelting of iron and other metallic ores.

It must not be inferred by the student that when wood and coal are destructively distilled, and carbonaceous material like charcoal and coke are left behind, *all* of the carbon of the wood or coal is thus left; most of the gaseous compounds also contain carbon, the principal one being methane, CH_4. Hydrogen gas is the other main constituent of coal gas.

For convenience, fuels are divided into three classes, solid, liquid, and gaseous.

Coal.—The principal *solid fuels* are coal, wood, and charcoal, coal being by far the most important industrially (Fig. 15). The *formation of coal* in the earth is a very interesting geological story, although it is still imperfectly known. At a time when the climate of the earth was very much warmer and more humid than it is now, when tropical jungles and swamps were prevalent over most of the United States, vegetation grew rapidly, died, and decayed rapidly. Tree ferns were one of the principal swamp plants of the time; as they fell they decayed in the swamps under water, were buried by other falling trees, continued to decay and to be buried deeper and deeper, were subjected thus to decomposition under hot swamps, and finally in the course of time had formed coal.

Depending on how far this decomposition had proceeded, we have various grades and kinds of coal, from lignite, one of the first stages, through soft coal to hard coal, which is the highest stage of this decomposition.
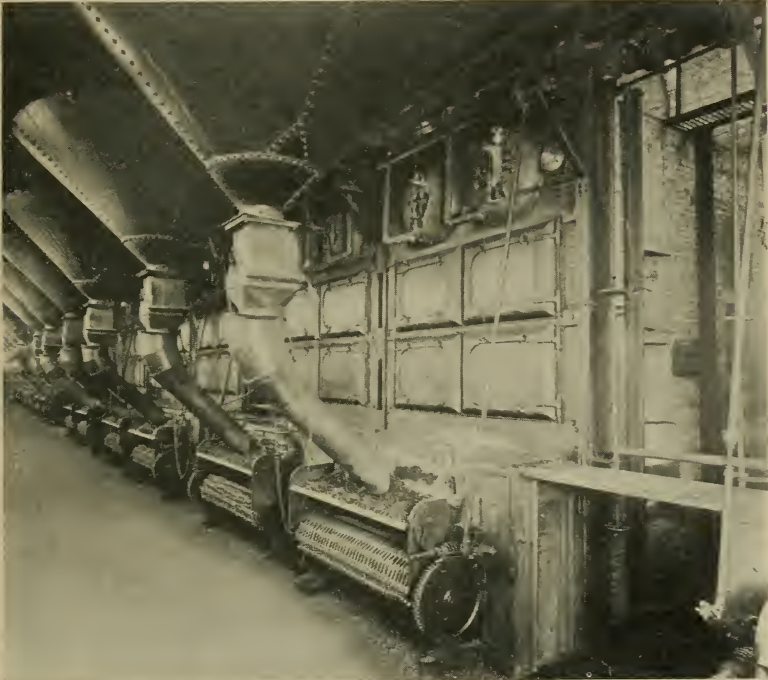
Fig. 15.—Battery of mechanical stokers. The coal flows down the chutes from bins and is fed into the grates in an even, steady stream, resulting in much more effective combustion than when it is hand-fired.

Peat.—Coal is probably not being formed anywhere on earth at the present time; the proper climatic conditions and kinds of plants are no longer present. But we do have decay of vegetable matter along such similar lines that we believe it is comparable to the first stages of coal formation. Thus in our peat bogs, the grasses and

mosses are dying down every year, being buried under new growth, and slowly decaying under water. The top layers of peat have a composition similar to that of the plants; the lower layers approach more nearly that of lignite. Hence we can say in a general way that coal has

Fig. 16.—The analysis of coal. A carefully weighed sample of coal is placed in the crucibles. These are placed in the small furnace and roasted with the covers on to determine the amount of volatile gases. They are then ignited with the covers off, to determine the amount of carbon and of ash. (Copyright, Underwood and Underwood, N. Y.)

passed through the following stages: live plants, dead plants, peat, lignite, soft coal, hard coal.

This process is shown also in the chemical analysis of these materials (Fig. 16). An analysis usually shows the amount of gaseous material which is driven off by roasting, as described above, the amount of carbon left behind

as coke, and the amount of ash. The following table presents the analyses of wood, peat, and coal:

TABLE III

Composition of Various Solid Fuels

	Gas per cent	Uncombined carbon per cent	Ash per cent
Wood..	50–70	20–35	3– 8
Peat..	40–60	20–40	10–30
Soft Coal..................................	20–35	50–75	5–15
Hard Coal.................................	5–10	80–90	5–10

It will be seen that as we go from wood through peat, to soft and to hard coal, there is a rapid increase in the amount of carbon, and a decrease in the volatile gases.

Fig. 17.—A California oil field. In order to speed up the production from a given claim, a great many wells are driven into the same oil beds. This is a waste of capital, however, since the beds are more quickly exhausted, and then the derricks have to be moved to a new location. (Courtesy of U. S. Geological Survey.)

And this is really the sum and substance of the chemistry of coal formation.

Alcohol.—Our common *liquid fuels* are practically limited to alcohol and to the petroleum products. Alcohol

is produced when yeast ferments sugar. The alcohol is distilled off and concentrated, and used for many purposes, including burning as a fuel in lamps. However, in this country it cannot be produced cheaply enough to be used very extensively in this way. Starch of potatoes and grains is changed to sugar and then to alcohol.

Petroleum.—Petroleum is found in enormous quantities in various parts of the world, especially in the United States, Russia, Roumania, and Mexico. The world's production of crude oil is about 460,000,000 barrels a year. Holes are drilled down into the oil beds and the oil pumped up like water (Figs. 17–18). In some cases the oil is under such tremendous pressure that it spouts out of the drill hole to a considerable height. This kind of well is called a "gusher." If it accidentally takes fire, it is very difficult to extinguish. The oil is often a black, disagreeable smelling liquid. As discussed in the last chapter, it is separated into several constituents by distillation. *Gasolene* finds most of its usefulness as the fuel in motors, while *kerosene* is mostly used in stoves or under boilers for steam production.

Fuel Gases.—The *gaseous fuels* that are produced and used to-day are numerous. The leading ones, of course, are the *coal gas* discussed above, and *natural gas*. The latter occurs in huge pockets beneath the ground in the oil regions of the world. It consists mostly of methane, CH_4. It is ready for burning without any purification, being very extensively used in the smelting industries, glass works, and for household use in cities. In small towns, where a factory for making coal gas would not pay, a gas called *water gas* is made, by passing steam over red-hot coke. The product formed is a mixture of two gases, carbon monoxide and hydrogen:

$$\underset{\text{coke}}{C} + \underset{\text{steam}}{H_2O} = \underset{\text{carbon monoxide}}{CO} + \underset{\text{hydrogen}}{2H}$$
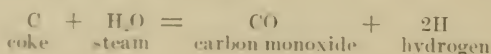
FIG. 18.—An oil tank on fire. Such fires are most spectacular, since there are huge quantities of oil involved, and since the poor access of air causes the production of enormous masses of black smoke. (Courtesy of U. S. Geological Survey.)

This carbon monoxide is not carbon dioxide, it should be noted. It will be taken up in the next chapter. *Acetylene* is another common gas for illuminating purposes. It is produced when calcium carbide is acted upon by water:

$$CaC_2 + 2H_2O = C_2H_2 + Ca(OH)_2$$
$$\text{calcium carbide} \quad \text{water} \quad \text{acetylene} \quad \text{slaked lime}$$
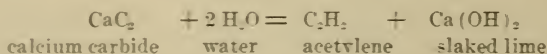
Acetylene generators are commonly installed in farm houses. The gas gives a brilliant white light, if it is provided with a proper burner, which admits the right amount of air for complete combustion. This gas is very much used in bicycle and automobile lamps. It is also used like hydrogen in the oxy-acetylene torch for welding and cutting steel (Fig. 19). Other gaseous fuels of less importance are also manufactured, but they need not be mentioned here.

Fire Extinguishers.—We have seen that a substance will not burn until it has been heated to its kindling temperature. Conversely, if a burning substance is cooled below its kindling temperature it will cease to burn. This principle is the basis of the use of water in *extinguishing fires*. The water thrown on a piece of wood keeps the temperature of the latter down, as the heat of the fires is used up in driving out the water, and if sufficient water has been used, the fire will be extinguished by the time the wood is dried out and raised to the kindling temperature. This is the reason why wet wood lights so slowly in a bonfire; the water must be driven out before the wood can be raised to the kindling temperature.

Another principle utilized in extinguishing fires is that, if oxygen (that is, air) can be excluded from the fire, the latter will have to die out, as combustion is not possible without oxygen. The air is excluded in a number of ways. In the first place, when water is thrown on a

Fig. 19.—The oxy-acetylene torch, used in cutting steel plates in the ship-building industry. Oxygen and acetylene are contained in the two tanks under pressure. The two gases mix and burn at the nozzle, forming a pencil of extremely hot flame which melts a cut through the steel. (Courtesy of Committee on Public Information.)

blaze, it not only cools down the materials, but the large volume of steam produced forces away the air, and thus aids in checking the fire. The commonest method of excluding air from a fire is to throw around the base of the fire large volumes of some gas that will not burn. This is the so-called chemical fire extinguisher. The usual gas that is used is carbon dioxide. We have seen above that this gas is formed when carbon burns; it is also formed when any acid acts on a carbonate, as we shall see in the next chapter. The apparatus used as shown in figure 20 consists of a tank, with a hose and nozzle attached. The tank is filled with a concentrated solution of baking soda (sodium bicarbonate); a n d suspended from the top is a bottle of sulfuric acid. As long as the two are kept separate there is no action; when the tank is inverted, however, the acid comes in contact with the carbonate solution, and carbon dioxide gas is generated and flows out of the hose, together with the water in a fine spray. The stream of gas and water is played around the base of the fire to exclude the air, and thus extinguish the flames. The reaction in the fire extinguisher is as follows:

FIG. 20.—Chemical fire extinguisher. When the machine is inverted, a cup of sulfuric acid is emptied into a reservoir of sodium bicarbonate solution. A large volume of carbon dioxide is generated, which issues from the hose mixed with the liquid, and can be sprayed around the base of the fire.

$$2\,NaHCO_3 \quad + \quad H_2SO_4 \quad = \quad Na_2SO_4 \quad + H_2O \quad + \quad CO_2$$
sodium bicarbonate   sulfuric acid   sodium sulfate   water   carbon dioxide

After using, the apparatus has to be cleaned and replenished. Fire department wagons of cities carry steel cylinders containing large volumes of carbon dioxide under pressure. Simply opening a valve liberates a jet of the gas from the hose attached. The throwing of a blanket on a burning object stops the fire by excluding the air in a similar manner.

QUESTIONS

1. What is oxidation?
2. Why is combustion important?
3. What are our two most important fuel elements? What product is formed by the combustion of each?
4. Name fuels that contain:
    (a) Carbon as an element.
    (b) Hydrogen as an element.
    (c) Carbon and hydrogen as compounds.
5. What is the kindling temperature of a substance? Explain why coal cannot be lighted by a match.
6. What is a flame?
7. Describe the different steps involved in the burning of a piece of wood.
8. Describe how coal is destructively distilled, and name the various products obtained from it.
9. Why is soft coal used for making coal gas instead of hard coal?
10. Make a list of all the fuels that you know, grouping them as solid, liquid and gaseous, and stating after each one whether it contains carbon and hydrogen as elements or compounds (get instructor's aid where needed).
11. Describe the successive stages in the formation of coal. What has been the fundamental chemical change here?
12. State the two general principles utilized in extinguishing fires.
13. How do the " chemical fire extinguishers " work?
14. Write an equation showing the reaction when acetylene burns.

LABORATORY EXPERIMENTS

17. **To Prove That Carbon Dioxide Is Produced by the Combustion of Food in the Body.**—By means of a glass tube, blow the breath for a minute or so into some clear limewater in a test tube. From what was learned of the chemistry of limewater in experiment 6, explain the results of the experiment. In what experiment was it proved that the combination of charcoal and oxygen produces carbon dioxide? Is the carbon which is burned in the body in the form of an element or of compounds?

18. **A Study of Flames.**—Make a drawing of the Bunsen burner, showing particularly the air holes at the base and the two principal portions of the flame. Where does the gas enter the burner? Where do gas and air first

mix? What is the result of closing the air holes? Prove to what the yellowness of this flame is due by holding a piece of cold porcelain in the flame. Why is this luminous flame wasteful of gas for heating purposes but more efficient for lighting? Of what does the inner cone of the non-luminous flame consist? Hold one end of a piece of glass tubing in this cone and light the gas issuing from the other end. Then hold the tube just above the inner cone. What do these tests prove concerning the progress of combustion in the flame? Ascertain whether a candle or an alcohol flame has the same structure as that of the Bunsen burner.

19. **The Destructive Distillation of Wood and Coal.**—Set up the apparatus shown in figure 12. Fill the test tube with hard-wood sawdust or small pieces of hard-wood, fix the cork in firmly, and carefully heat the test tube until the gases given off will ignite. Continue the heating until no further inflammable gas is given off. What are some of the constituents of this gas? How were they formed? Examine the residue in the test tube. Place it on a wire gauze and ignite it with a burner until it ceases to burn. Does it burn with much flame or does it glow like charcoal? Of what does the final incombustible residue consist? Repeat the above experiment, using soft coal. What is the residue in the test tube called in this case? What products are separated from coal gas commercially?

CHAPTER VI

CARBON

Occurrence of Carbon.—We have already discussed the element carbon to a considerable extent. We have seen that it occurs in the atmosphere as carbon dioxide in small amounts; that plants extract large quantities of it from the air to build up their bodies; that coal is principally carbon, especially hard coal; that when plant material is roasted without access of air, charcoal, which is practically all carbon, is left behind; that almost one-half of the dry matter of all plant and animal bodies is carbon; that it is the most important of our fuel elements. But even these things do not cover the usefulness of carbon; there are many more interesting chemical stories concerning it, which we will relate in this chapter.

Although a very common element in our daily life, carbon constitutes only one five-hundredth of the weight of the earth's crust. It forms more compounds than any other element, something like 100,000 of them being known at present. In one form (diamond) it is one of our most beautiful gems; in another form (soot) it is one of our blackest and grimiest of substances.

Forms of Carbon.—Carbon as an element occurs in three forms: diamond, graphite and charcoal. *Diamond* is pure, crystalline carbon; it has a brilliant lustre, and breaks up light into many colors. Most diamonds are white, while others are tinged with color due to mineral impurities. It is the hardest substance known, and hence furnishes a valuable means of grinding and cutting other hard substances. Broken and badly colored pieces are used in glass cutters, rock drills, and grindstones for other

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gems. We know that diamond is carbon because it burns
in oxygen to form nothing but carbon dioxide. Its origin
in the earth is not known exactly, but the carbon no doubt
had an animal or vegetable source; under certain con-
ditions of heat and pressure charcoal was formed, and this
underwent still further effects of heat and pressure, be-
came liquid carbon, and then crystallized into diamonds.
We believe this to be the process because it has been
imitated artificially. The brilliant French chemist, Henri
Moissan, not many years ago, dissolved pure charcoal in
molten iron, in an electric furnace, then plunged the
mass of white-hot iron into water. The tremendous pres-
sure of the contracting and cooling iron resulted in the
formation of a great many very small crystalline dia-
monds, which were found imbedded all through the iron.
These crystals are just as hard and as brilliant as the
natural gems, but as yet they cannot be made large enough
for jewelry purposes.

Graphite is also crystalline carbon, although it has a
different crystal form from that of diamond; is black or
grayish black, and very soft and greasy to the touch.
These latter qualities make it a valuable lubricant, espe-
cially when ground in oil. Another common use of
graphite is in lead pencils. Here it is mixed with clay;
the more clay that is used the harder the pencil. It is also
incorporated into paints for iron parts. Although
graphite occurs in fairly large quantities in Ceylon,
Siberia and the United States, a great deal of it is made
artificially, by fusing coke in an electric furnace, shown
in figure 21.

By the term *charcoal* here we mean any form of carbon
other than diamond and graphite. We have seen in the
last chapter that charcoal is produced when the gaseous
materials are driven out of wood; that coke is a form of
charcoal produced in a similar way from coal, and that

coal, especially hard coal, is mostly carbon, and hence we can think of it as being the charcoal produced in nature. Both our charcoal from wood and our coke from coal are light, porous substances, but we can imagine that if they were subjected to great grinding and pressure in the earth, they would come out of it in hard, compact form like coal. In other words, we can go to nature and find
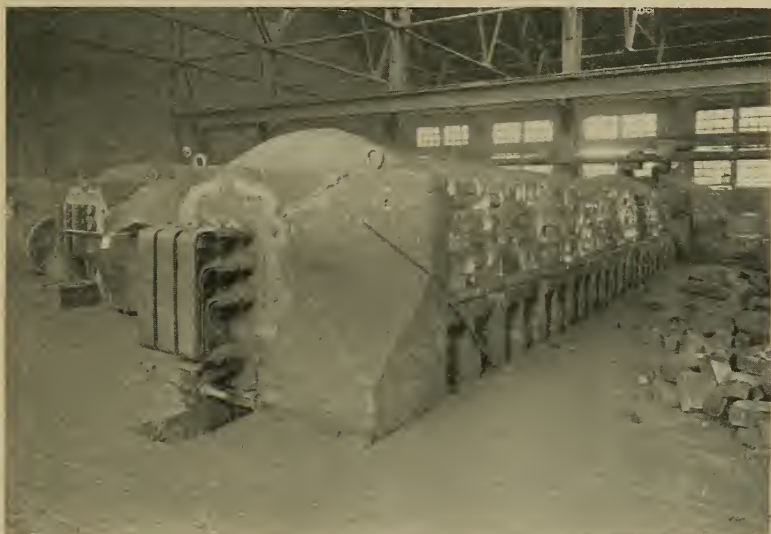
Fig. 21.—An electric furnace for making graphite. This is the only kind of furnace which develops a sufficiently high temperature for the formation of the graphite form of carbon.

all the stages of carbonaceous decomposition; starting with living plants, then dead plants, then peat, lignite, soft coal, hard coal, graphite, and end up with absolutely pure, crystalline carbon in the form of diamond. To be sure, we do not know that exactly the same process was involved here from start to finish; but the series at least illustrates that nature is provided with chemical processes for obtaining pure carbon from plant and animal tissues.

In all probability plant and animal tissues were subjected to a somewhat different treatment in other parts of the earth and at different times, and resulted in our beds of petroleum and pockets of natural gas. We shall have occasion to mention other deposits of minerals which prob-

ably had an origin in plant or animal materials, such as the nitrate beds of Chile, the phosphate rock of Florida, and the huge deposits of limestone all over the world. These facts indicate the close connection between the inorganic or mineral world and the organic or living world.

Kinds of C h a r c o a l.— Charcoal is given a special n a m e, according to its source. Thus, the charcoal from kerosene is called *lampblack;* f r o m bones, *bone-black;* from blood and other slaughter house residues, *a n i m a l charcoal;* from wood, *wood charcoal.*

FIG. 22.—Apparatus showing the power of charcoal to absorb gases. The test tube is filled with ammonia, hydrogen sulfide, or other gas, and then set down over a piece of fresh charcoal floating on the mercury. The charcoal absorbs the gas, creating a partial vacuum, drawing mercury up into the tube.

From w h a t e v e r source, charcoal is a jet-black, light, porous, soft, easily powdered substance, very resistant to the action of chemicals. Its most remarkable property is that of absorbing gases and coloring matter. A piece of fresh charcoal will absorb eighty times its volume of ammonia or other gas (Fig. 22). It is commonly thrown into vaults, cisterns, and refuse piles to absorb offensive odors. When powdered charcoal is shaken up with colored liquids and filtered, a great

deal of the coloring matter is removed. Great quantities of charcoal are used in this way in purifying cane and beet juices in sugar manufacture.

Battery Carbons.—When charcoal is powdered and pressed into solid bars it is given the trade name of carbon. As such it is used in many electrical instruments, such as one of the poles of batteries, the poles of arc lights and lanterns, in telephone and telegraph instruments, and the resistance elements in electric furnaces.

INORGANIC COMPOUNDS OF CARBON

Carbon unites readily with many other elements. In fact, there are many thousands of compounds of carbon. Therefore, in order to simplify the study of them, they are divided into two great groups, the *inorganic* and the *organic*. In general, the organic compounds are found in plants and animals, and the inorganic in the non-living world, although this distinction does not always hold.

Carbon Dioxide.—We have already mentioned two different combinations between *carbon* and *oxygen.* The ordinary one, *carbon dioxide,* is produced when carbon or a compound of carbon burns with plenty of oxygen present. Thus, it is formed during the burning of coal or charcoal in our stoves and furnaces, when wood is burned, when gasoline explodes, when coal gas, acetylene, alcohol and kerosene burn, and when food is burned in the animal body. As has been pointed out before, carbon dioxide is also formed when any acid acts on any carbonate. This is made use of in fire extinguishers (Fig. 20). The use of soda in cooking in connection with acids, as in sour milk and vinegar, is explained by the fact that the baking soda reacts with the acids to produce carbon dioxide, and this gas puffs up the dough and makes it light and porous. Baking powder is a mixture of an acid (tartaric or phosphoric) and baking soda. As long as they are dry there

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is no reaction; on mixing with wet dough, however, carbon dioxide is generated as in the case of soda and sour milk above. When yeast raises bread dough, it generates carbon dioxide gas; not by means of a carbonate, of course, but simply as a product of combustion, since yeast is a living organism and produces carbon dioxide when it consumes food, the same as any higher animal or plant. In the same way the effervescence or " fizz " of beer and wine is due to the carbon dioxide produced by the yeast during the fermentation. Large volumes of carbon dioxide can be stored in steel cylinders under pressure; these cylinders are sold to soda fountains and to manufacturers of soft drinks to " charge " their drinks with " fizz " (Fig. 23). These cylinders are also used in fire extinguishers, as has been mentioned.

The Carbon Cycle.—The story of carbon dioxide in nature is a rather remarkable one. We can narrate here in a few lines what required in nature millions of years. In the earlier history of this globe, there was a great deal more carbon dioxide in the air than there is at present. Also, a far greater proportion of the earth was covered by the seas. These seas were alive with plant and animal life. The plants absorbed the carbon dioxide of the air. Minute sea-animals ate the plants and stored the carbon in their own bodies, putting some of it into their skeletons as calcium carbonate. When the animals died, the soft body parts decayed and the skeletons dropped to the bottom of the seas. After long series of
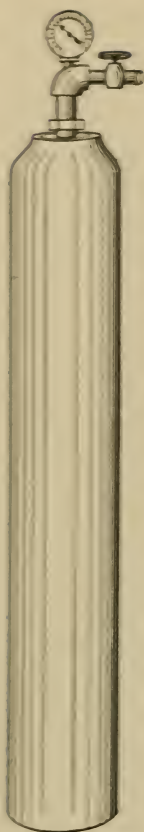
Fig. 23.—Steel cylinder containing liquid carbon dioxide. Many other gases, such as oxygen, hydrogen, sulfur dioxide, and chlorine, are stored in such cylinders, in which they can be kept indefinitely and used wherever and whenever wished. See also figure 8.

years, the accumulations of these skeletons amounted to
beds of calcium carbonate hundreds of feet thick in some
places. These beds are our limestone, marble and chalk
beds of to-day. Wherever calcium carbonate in one of
these forms is found, it means that that region was once
the bottom of a sea. Thus the carbon dioxide of the air
passed through a plant life, then an animal life, then spent
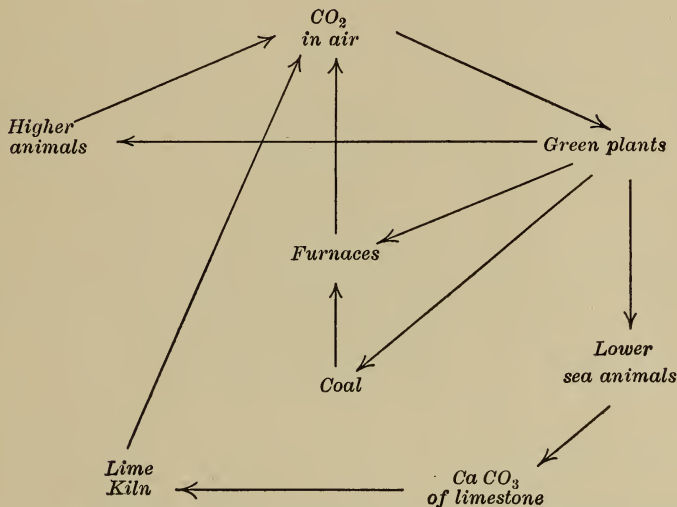
FIG. 24.—Diagram showing the cycles of carbon. The carbon of the carbon dioxide in the
air can be traced through the plant world, the animal world, the mineral world, to the
industrial world, and thence to the air again. Shorter cycles can also be traced.

ages and ages in the ground. Now it is mined as lime-
stone; and when we "burn" limestone to make quicklime,
the carbon dioxide is driven off again into the air, thus
completing the "cycle of carbon." There is an enormous
amount of carbon dioxide stored up in the limestone of
the earth, as this mineral constitutes whole mountains,
and beds of it hundreds of square miles in extent are
found in all parts of the world. Figure 24 shows this
story of carbon in a diagrammatic way.

The fact that when limestone is roasted to make quick-lime carbon dioxide is given off, provides us with a ready means of detecting carbon dioxide. When the quicklime is slaked in water, we get the so-called limewater, which is a perfectly clear solution. When air, or any gas containing carbon dioxide, is passed into this solution it becomes cloudy because of the re-formation of calcium carbonate.

The other oxide of carbon is called *carbon monoxide*. It is produced when carbon burns in an insufficient amount of oxygen; that is, only half as much oxygen combines with the carbon as when the dioxide is formed. The two equations involved bring this out:

$$C \; + \;\;\; O \; = \;\; CO \quad \text{(carbon monoxide)}$$
$$C \; + \; 2\,O \; = \;\; CO_2 \quad \text{(carbon dioxide)}$$
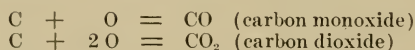
Since the carbon in the first instance is really only half burned, we might expect that carbon monoxide would undergo further oxidation; and indeed it does burn very readily:

$$CO \; + \; O \; = \; CO_2$$

On the other hand, when carbon dioxide is passed through red-hot carbon, the latter takes away a part of the oxygen from the dioxide, and carbon monoxide results:

$$\underset{\text{carbon dioxide}}{CO_2} \; + \; \underset{\text{red-hot coal}}{C} \; = \; \underset{\text{carbon monoxide}}{2\,CO}$$
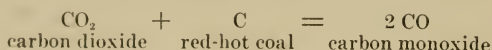
It is in this way that hard coal stoves are always producing carbon monoxide, which burns with a pale-blue flame at the top of the pile of coal (Fig. 25). If the draft is too suddenly shut off from a very hot fire, the monoxide escapes into the room and often suffocates sleeping people. For this oxide of carbon is extremely poisonous, acting on the red blood corpuscles in such a way that they can no longer absorb oxygen from the lungs. It

cannot be detected by its odor, for it has none; but it is usually accompanied by sulphurous odors from the coal, which indicate when the stove is not working properly. Carbon monoxide is also emitted from volcanoes, and is one of the two gases in water gas (p. 54).

Carborundum.—When sand, which is silicon dioxide,
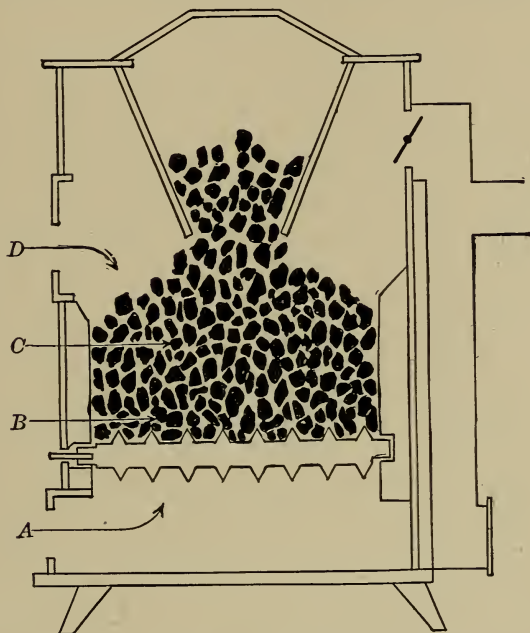
FIG. 25.—Diagrammatic section of a hard-coal stove. Fresh air enters at A; the oxygen of the air combines with the carbon of the coal at B to form carbon dioxide; the carbon dioxide comes in contact with hot carbon at C, and is reduced to carbon monoxide. This, being an inflammable gas, burns with a pale-blue flame at the surface of the pile of coal by means of fresh oxygen entering at D. If no air is allowed to enter at D, the unburned gas passes out the chimney, or escapes into the room and poisons the occupants.

SiO_2, is fused in the electric furnace (Fig. 21), at a very high temperature with coke, part of the carbon combines with the silicon to form *silicon carbide,* or *carborundum:*

$$SiO_2 + 3C = SiC + 2CO$$
sand   coke   silicon carbide   carbon monoxide
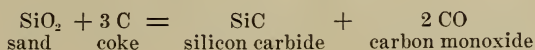
This carborundum comes out as brilliant, beautiful, bluish-black crystals; they are extremely hard, and are used to make carborundum grinding stones and wheels.

Carbide.—Another common carbide is one which is a compound of calcium and carbon. It is called *calcium carbide*, or, very commonly, simply carbide. It is produced by fusing quicklime and coke in the electric furnace. It is a very familiar substance, being used for generating acetylene, an inflammable gas. This has been mentioned before on page 56. The gas is used for illuminating houses or other buildings, and in bicycle and automobile lamps. In the latter cases it is dissolved in acetone and stored in steel cylinders.
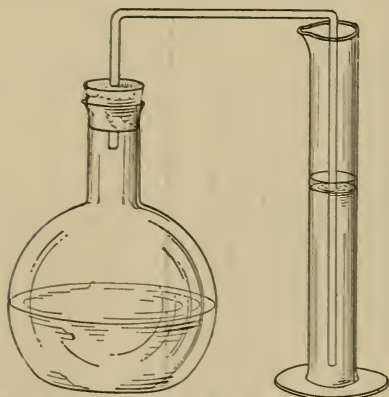
Fig. 26.—Apparatus used in the study of yeast fermentation.

Carbon Bisulfide.—Carbon combines with sulfur to form *carbon bisulfide*. This is a disagreeable smelling liquid, the vapors of which are highly inflammable and very poisonous. The latter property enables it to be used as a poison for insects, gophers, prairie dogs, and vermin. When cotton soaked in the liquid is dropped into the burrow, the heavy vapors flow down the hole and kill the occupants. Carbon bisulfide is used extensively in vulcanizing rubber, making artificial silk, and dissolving sulfur.

ORGANIC COMPOUNDS OF CARBON

In taking up now some of the compounds of carbon with nitrogen, hydrogen and oxygen, we are confronted with an enormous task. Most of the substances in animal

and plant bodies are made up of these elements. The study of organic compounds alone forms such a huge branch of chemistry that it is almost a science in itself. It is called *organic chemistry,* because in the early history of these compounds it was thought that they were found only in plant and animal bodies and were thus the products of organic life. Since then, however, thousands of such compounds have been made artificially that have never been found in plants or animals. Among these are many drugs, most of our dyes, explosives, some flavoring substances, disinfectants, artificial silk, and photographic chemicals. Then, too, many natural compounds have been made artificially by organic chemists; rubber, indigo, and the flavoring substance of vanilla and of many fruits are among these.

Then, too, organic chemistry is of vital assistance in many manufactures, where the raw materials of nature are prepared for human food, clothing, shelter, and amusement. Thus organic chemists control the processes of making granulated sugar from cane and beets; of curing tobacco and blending the varieties to make various products; of malting and brewing; of making corn syrup out of cornstarch; of making various stock feeds as by-products from flour and cereal mills; of making artificial silk; of distilling turpentine, petroleum, and perfumes; of making medicines, stock dips, insect poisons, disinfectants; of canning fruits and vegetables; of manufacturing candy, dairy products, soap, candles; of the packing house industries, of the paint, oil and varnish factories.

To state it all very briefly, organic chemistry has to do with all substances and materials that contain carbon, and with all the processes through which man puts them to fit them for his needs. Hence, if we can but begin to realize the universal abundance of the compounds of carbon, we will appreciate the necessity of knowing **at**

least the A B C's of organic chemistry. Therefore, let us consider for a while a few typical classes of the compounds of carbon.

Hydrocarbons.—First of all there are the compounds of carbon with hydrogen alone. They are therefore called the *hydrocarbons. Methane,* CH_4, is the simplest. It is the "fire damp" met with by miners, causing explosions in mines. It is also called "marsh gas," as it is the gas that bubbles up from the bottom of stagnant pools in hot weather. It constitutes the bulk of natural gas, and a large portion of coal gas.

All the petroleum products are hydrocarbons; thus *gasolene* is mainly C_6H_{14}; *benzine,* C_8H_{18}, and *kerosene,* $C_{10}H_{22}$. *Acetylene,* C_2H_2, is another. There are two important hydrocarbons obtained as by-products in the making of coal gas: they are *benzol,* C_6H_6, and *toluol,* C_7H_8. It will be noticed that all of these compounds are highly inflammable; they constitute our most important liquid and gaseous fuels. This is to be expected, as *they are compounds of the two great fuel elements, carbon and hydrogen.* Theoretically, therefore, when gasolene or any other hydrocarbon is burned completely, carbon dioxide and water should be the only products formed; thus:

$$C_6H_{14} + 13\,O = 6\,CO_2 + H_2O$$
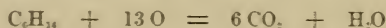
These should have no particular odor. The exhaust gases from a gasolene engine do have an odor, however, for at least two reasons: In the first place, the combustion mixture of gasolene vapor and air is seldom perfect; that is, there are not exactly 13 atoms of oxygen for each molecule of gasolene, as is demanded by the above equation, and hence other products are formed which do have odor. In the second place, some of the lubricating oil in the cylinders is decomposed and charred by the heat, and results in odorous compounds and the "cylinder carbon."

Paraffine, a solid, non-volatile petroleum product, has a molecule represented mainly by the formula $C_{24}H_{50}$; it will be seen that if this molecule could be decomposed into three or four parts, some of these parts would have about the same number of carbon atoms as gasolene. This is actually done in the process called "cracking." Within the last few years chemists have been able to obtain considerably more gasolene from the crude oil by cracking the heavier constituents.

Aldehydes.—The next simplest carbon compounds are those where oxygen enters into the molecule along with the carbon and hydrogen. There are many different classes of such compounds, each of which has many individual members. The first class is the *aldehydes,* the simplest member of which is *formaldehyde,* HCHO. Formaldehyde is an irritating poisonous gas. When it is dissolved in water to make a 40 per cent solution, it goes by the commercial name of "formalin." This finds very important use in treating potato seed for scab, in disinfecting and fumigating rooms, barns, chicken houses, etc., and in preserving animal specimens for classroom study. There is another very interesting fact concerning formaldehyde: it probably is the first product formed when green leaves manufacture sugar in the sunlight from carbon dioxide and water. The equation involved is:

$$CO_2 \ + \ H_2O \ = \ HCHO \ + \ 2\,O$$
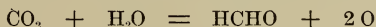
As fast as the formaldehyde is formed, six molecules of it combine together to form one molecule of glucose sugar:

$$\underset{\text{formaldehyde}}{6\,HCHO} \quad \underset{\text{sugar}}{= C_6H_{12}O_6}$$
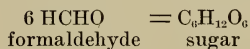
Since the plant then uses this glucose to make all the other substances in its body, we can consider that for-

maldehyde is the basic material from which all other plant constituents are made. Another common aldehyde is *benzaldehyde*, C_6H_5CHO. This gives the characteristic odor and flavor to peach and cherry stones; artificial almond, peach, and cherry flavoring extracts consist essentially of benzaldehyde.

Alcohols.—The alcohols form another very important class of organic compounds. Among the great many members of this class we can mention but a few. *Methyl alcohol*, or wood alcohol, or wood spirit, as it is variously called, is CH_3OH. It is produced, as we have learned before, when wood is roasted or distilled. It finds very great use in the paint and varnish industries, also as a fuel in small burners and in making formaldehyde. It is a strong poison. *Ethyl alcohol*, or grain alcohol, is C_2H_5OH. It is produced almost exclusively by the fermentation of sugar by yeast, according to the following equation:

$$C_6H_{12}O_6 = 2 C_2H_5OH + 2 CO_2$$
$$\text{dextrose} \quad \text{alcohol} \quad \text{carbon dioxide}$$
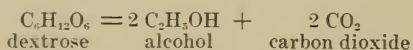
And since yeast cells are constantly present in the air and on fruit, whenever fruit juices or other sugary solutions are exposed to the air alcoholic fermentation sets in. Wine and hard cider are made by this natural means of infection. In making alcoholic beverages from grains, the starch of the grain must first be converted into sugar. This is done by sprouting the grains, then mashing up the sprouted grain with water. Yeast is then added, and the fermentation allowed to proceed at the proper temperature. About 5 per cent of alcohol is usually produced. If a beverage containing a much higher content of alcohol is wanted, the fermented mash is distilled, as was described a few chapters back. The first distillate is from

35 per cent to 45 per cent alcohol. This constitutes whiskey, brandy, etc., after it has aged in oak barrels for the proper length of time. If pure alcohol is wanted, two or three redistillations are employed, or else a special kind of single distillation. This gives a 95 per cent alcohol, which is the most concentrated that is ordinarily found on the market. It is used in medicine, as a solvent for certain paints and chemicals, and in many other arts and industries. Since it is subject to a heavy government tax, however, it is too costly for most purposes. When it is "denatured," that is, made unfit for drinking purposes by the addition of some poisonous or bad-smelling substance, the tax is removed, and this cheapens the cost to the consumer considerably. This denatured alcohol now finds very extensive use in the industries and as a fuel in small lamps. Alcohol has a pleasant, aromatic odor, and a burning taste; taken into the body in small quantities it is a stimulant, but in larger quantities is a poison.

Glycerin.—Another substance which is an alcohol, but which has entirely different properties from the two above, is *glycerin*. This is a constituent of all fats and oils of animal or plant origin. It is a by-product in soap manufacture. It is a white, syrupy liquid, used considerably in skin ointments. Its greatest use, however, is in the manufacture of nitroglycerin, the explosive, which is used alone or is absorbed on fuller's earth and pressed into sticks as dynamite.

Acids.—The *organic acids* form another common class of substances, found very extensively in plants and animals of all kinds. *Formic acid*. HCOOH, is the irritating substance in ants and stinging nettles. *Acetic acid,* CH_3COOH, is the acid in vinegar. It is produced by a bacterium which, in feeding on the alcohol of hard cider

and other fermented fruit juices, converts it into acetic
acid. The following equation shows this:

$$C_2H_6O \ + \ 2\,O \ = \ C_2H_4O_2 \ + \ H_2O$$
$$\text{alcohol} \qquad \text{oxygen} \qquad \text{acetic acid} \qquad \text{water}$$
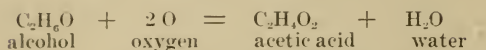
This is very evidently an oxidation process. Everyone
who has made vinegar knows that the fermented cider
must be kept in casks which are not over one-half or three-
fourths full, and that the casks must be kept open. We
can now understand that this is because plenty of oxygen
must be available to the vinegar bacteria for them to
carry on their work. Acetic acid is also formed in silage
and sauerkraut, and gives them their acid odor and taste.

Butyric acid is formed when butter becomes rancid,
and gives the latter its rancid odor. *Stearic, palmitic,*
and *oleic* acids are the main constituents in most fats and
oils; they will be mentioned later under that heading.
Lactic acid is the sour milk acid, produced when the lactic
acid bacterium ferments the milk sugar. It is also present
in silage. *Malic acid* is the principal acid in apples, *citric
acid* in lemons and gooseberries, *tartaric acid* in grapes,
tannic acid or tannin in tea and in oak and hemlock bark,
and *oxalic acid* in rhubarb.

Mention might be made here of that series of very
delightful compounds which give flowers their odors and
fruits their flavors, and from which we make our per-
fumes and incense oils. They are given the general name
of *esters and volatile oils.* That they serve a useful pur-
pose in nature cannot be doubted; they attract insects to
flowers and thus help pollinate them; they give flavor and
palatability to fruits and vegetables; spices help preserve
foods; oil of eucalyptus, oil of turpentine, oil of cedar,
rosin, wintergreen, peppermint, vanilla, oil of lemon, of
orange, of cloves, of roses, are all useful. The world is
more interesting and pleasant because of these.

Among the more complex compounds of carbon, hydrogen, and oxygen are the carbohydrates and the fats and oils. As these are of primary importance in human and animal foods and in many agricultural occupations, they will be accorded more space and detail in Chapter XV. The same will also be done with the compounds in which nitrogen also enters into the composition, with the exception of the following two compounds:

Hydrocyanic Acid.—When one atom each of carbon, nitrogen, and hydrogen unite, there is formed the very simple substance of the formula HCN. Its chemical name is *hydrocyanic acid;* its common name is *prussic acid.* From its name it will be recognized as the well-known deadly poison. In fact, it is one of the most deadly poisons known. In nature it occurs in the young plants of sorghum cane, and often kills cattle which eat it; in bitter almond oil; in small amounts in peach and cherry stones; in flaxseed, where it is poisonous in the oilmeals of cattle feed unless they have been heated in the process of manufacturing, which destroys the poison. Hydrocyanic acid is obtained from potassium cyanide, or prussiate of potash; it is used in fumigating buildings, greenhouses, orchard trees, and flour mills to rid them of insects (Fig. 27).

Cyanamide.—A compound of carbon with calcium and nitrogen is *calcium cyanamide,* $CaCn_2$, a substance unknown a few years ago, but one that is rapidly coming into use as a nitrogenous fertilizer.

Chloroform and *iodoform* are two other compounds which do not fit into any of the above classes of organic compounds. The element chlorine enters into the first, and iodine in the other substance. Thus chloroform is $CHCl_3$, and iodoform is CHI_3. Chloroform is the familiar anæsthetic, and iodoform is a disinfectant commonly used in hospitals.

Summary of the Chemistry of Carbon.—In glancing back

through the contents of this chapter, the student may express surprise at the great variety of subject matter grouped together under the heading "carbon"; there is diamond and charcoal; the poisonous gas, carbon monoxide; fire extinguishers; the "fizz" of soft drinks; limestone beds; petroleum and natural gas; carborundum grindstones; gopher poisons; prussic acid; sugars; per-
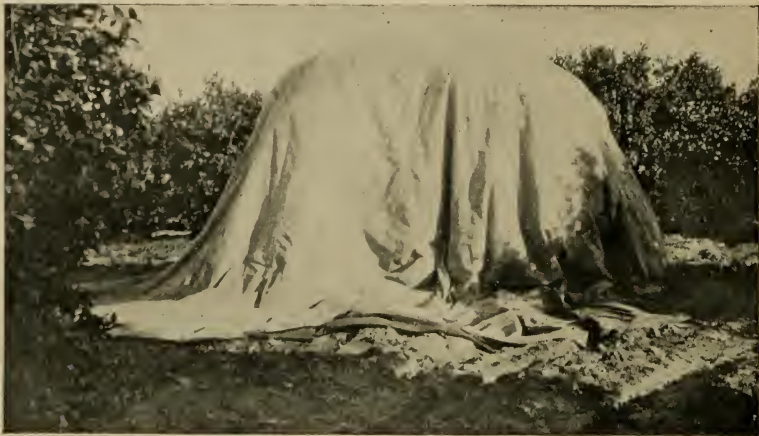
Fig. 27.—Tent over orange trees for fumigating with hydrocyanic acid gas to kill scale insects. This method is more common in California than elsewhere. The gas is either generated by dropping sodium cyanide into sulfuric acid, or liquid hydrocyanic acid is purchased direct. (From Farmers' Bul. 923, U. S. Dept. Agric.)

fumes; fats; alcohol; the acid of silage; anæsthetics. It is well if such an impression has been received by the student, since it will serve to emphasize that carbon has more or less rightly been termed the "universal element."

Carbon is very abundant in the earth in the form of many useful minerals, oils, and coals; it is even more abundant in all animal and plant bodies. Indeed, carbon is the main connecting link between the living portion of the world and the non-living portion, for the great masses of non-living forms of carbon in the earth no doubt had

their origin, some time in the past, in the remains of dead animals and plants. Then, as we take coal from the ground, burn it, throw the carbon dioxide produced into the air, and grow plants which absorb this carbon dioxide again, are we not really bringing the dead carbon back to a living form again? This is done in still another way. We saw that the minute sea animals eat plants, which get their carbon from the air; the animals store some of this carbon in their skeleton in the form of calcium carbonate; the animals die, the skeletons drop to the bottom of the ocean and form limestone; when we burn quicklime we drive the carbon dioxide out of the limestone, and return it to the air from whence it came in the first place. This carbon, then, has in turn served the life of the sea, the rocks of the earth, and the industries of civilized man. It may next be absorbed by a potato plant, and then eaten by man, and thus actually form a part of man's body itself. If atoms of carbon could speak, there would be some marvellous tales to be told, covering untold millions of years in time, and involving existence in air, rocks, and water, the strange plants and the freak animals of pre-historic times, in peat bogs, in the white-hot furnaces of the earth, and finally, perhaps, in a diamond in the crown of a king.

These more or less fanciful facts are mentioned in order to illustrate the multitude of important functions that carbon and its compounds serve in the world. It is not only of fundamental necessity in all living tissues, but it is also intimately concerned with the inorganic or mineral world.

The table on page 80 constitutes a summary of the chemistry of carbon and its compounds as we have discussed it. A thorough study of this table will not only help to fix some of the facts in the mind, but it will aid in establishing the relation between the compounds and in acquiring a mental picture of the whole carbon world.

TABLE IV
Summary of the Chemistry of Carbon

	Formula	Where found or how obtained	Uses
As the element:			
Charcoal, lampblack, coke	C	By roasting coal, wood, bones, etc.	Deodorant; decolorizer; printers' ink; fuel.
Graphite..........	C	In mines, and made artificially	Lubricant; electrical implements.
Diamond.........	C	In mines, and made artificially	Jewelry; cutting glass and stone.
Combined with oxygen:			
Carbon monoxide....	CO	In coal stoves; in coal gas	Fuel.
Carbon dioxide......	CO_2	In air; by burning of carbon	Fire extinguishers; soft drinks; raising bread.
Combined with hydrogen:			
Methane, or marsh gas	CH_4	In mine gas; in coal gas	Fuel.
Gasolene............	C_6H_{14}	From petroleum	Fuel.
Benzine............	C_8H_{18}	From petroleum	Fuel.
Kerosene	$C_{10}H_{22}$	From petroleum	Fuel.
Acetylene..........	C_2H_2	Action of calcium carbide on water	Illuminant and fuel.
Benzol, or benzine ...	C_6H_6	From coal tar	Fuel; solvent; explosives.
Toluol, or toluene....	C_7H_8	From coal tar	Fuel; solvent; explosives.
Naphthalene........	$C_{10}H_8$	From coal tar	Moth balls.
Combined with carbon and oxygen:			
Formaldehyde.......	CH_2O	Oxidation of wood alcohol; produced in plants to form sugar	Disinfectant.
Benzaldehyde.......	C_6H_5CHO	Peach and cherry stones	Artificial cherry flavor.
Wood alcohol.......	CH_4O	By distilling wood	Fuel; varnishes.
Grain alcohol........	C_2H_6O	By fermenting starch or sugar	Fuel; preservative; solvent; stimulant.
Glycerin............	$C_3H_6O_3$	In manufacture of soap	In making nitroglycerin; cosmetic.
Acetic acid.........	$C_2H_4O_2$	In vinegar and silage	
Butyric acid........	$C_4H_8O_2$	In butter	
Lactic acid.........	$C_3H_6O_3$	In sour milk and silage	
Tannin............	....	In oak and hemlock bark; in tea	In making ink.
Starch.............	$C_6H_{10}O_5$	In most seeds and tubers	Food; laundering.
Dextrose sugar ..	$C_6H_{12}O_6$	In corn syrup; in many fruits	Food.
Cane sugar........	$C_{12}H_{22}O_{11}$	In sugar cane; sugar beet; maple sap	Food.
Miscellaneous compounds			
Carborundum.......	SiC	Made in electric furnace	Grindstones.
Calcium carbide....	CaC_2	Made in electric furnace	For making acetylene gas.
Carbon bisulfide....	CS_2	Made artificially	Poison; vulcanizing rubber.
Hydrocyanic acid....	HCN	In flaxseed; from potassium cyanide	Killing insects.
Chloroform	$CHCl_3$	Made artificially	Anæsthetic.
Iodoform........	CHI_3	Made artificially	Disinfectant.

QUESTIONS

1. How does carbon occur in the air? In the earth?
2. Name the stages involved in the formation of pure carbon in nature.
3. What uses have diamonds?
4. What is formed when graphite burns? Write an equation.
5. How are diamonds made artificially?

6. What is the principal use of graphite?

7. Where did petroleum and natural gas probably have their origin?

8. How is charcoal made? What are some of the by-products obtained?

9. How is coke made? What are some of the by-products obtained and for what are they used?

10. Describe how each of the oxides of carbon is produced, state when and how each can be produced from the other, and state what the importance of each is.

11. Give three important uses of carbon dioxide, and state how the gas is prepared or handled in each case.

12. How can one prove that there is carbon dioxide in the breath? How did it get there?

13. Beginning with coal, describe by what process we can get acetylene gas for our automobile lights.

14. State at least one use for each of the following: Carbon bisulfide, carborundum, calcium carbide, prussic acid, methane, wood alcohol, kerosene, formaldehyde, iodoform, glycerin, and acetic acid.

15. How is pure alcohol made, starting with barley grain?

16. How is denatured alcohol made, and what are its uses?

17. Give briefly the story of an atom of carbon, beginning with the carbon in limestone and ending up with diamond.

18. What has the science of organic chemistry done for our civilization?

LABORATORY EXPERIMENTS

20. To Show the Absorptive Power of Charcoal.—Roast several small pieces of charcoal in a covered dish, to drive out all absorbed gases. Then perform the following tests.

(a) Place several pieces of charcoal in a test tube containing hydrogen sulfide water, or if this is not available, dilute ammonia water, shake for a few minutes, then filter, and test for the odor of hydrogen sulfide. Compare the amount of the odor with that of the water before treatment. In what other way can the absorption of gases by charcoal be shown? (Fig. 22.)

(b) Repeat the above experiment using a solution of brown sugar or of litmus, boiling for a moment before filtering. What uses for charcoal are suggested by the results of (a) and (b)?

21. The Preparation and Study of Carbon Dioxide.—Set up the apparatus shown in figure 11. Place in the generating flask several pieces of marble or limestone the size of peas. Pour dilute hydrochloric acid down the funnel. After about a minute collect the gas as was done in experiment 12 and perform all the tests on it that were performed on hydrogen and on oxygen and record the results. Write the equation. Describe carbon dioxide, as to its color, inflammability, and density compared to air. Pass some of the gas into limewater. Where was this experiment performed before? Close up all but one of the air-holes of a lighted Bunsen burner, and pass the jet of carbon dioxide into this hole. Explain the results. Place some baking soda ($NaHCO_3$) in a test tube, and add vinegar to it; then hold a lighted match over the mouth of the tube. How do fire extinguishers work?

22. To Demonstrate the Presence of Carbon in Various Substances.—Heat separately in porcelain crucibles small amounts of the following substances: Sugar, cotton, wool, sand, salt, soap, powdered lime-

6

stone, and meat. What does blackening indicate? What kinds of substances will not blacken on heating?

23. The Fermentation of Dextrose by Yeast.— (To be done by instructor.) Place 150 c.c. of glucose syrup in a liter flask with 500 c.c. of water. Add the following salts to supply mineral nutrient for the yeast; 2.0 g. potassium phosphate, 0.025 g. calcium chloride, 0.025 g. magnesium sulfate and 2.0 g. ammonium sulfate. Mix a half cake of compressed yeast with water to form a thin paste and add this to the flask. Fit the flask with a stopper and bent glass tube as shown in figure 26. Put clear limewater in the cylinder with a thin layer of kerosene on the surface. Stand the apparatus in a warm place for several days, noting the appearance of the limewater at frequent intervals. Explain the results. What is the purpose of the kerosene? After the fermentation has proceeded several days, decant the clear portion of the liquid in the flask into a distilling flask and distill off about 200 c.c. Redistill this 200 c.c. to 100 c.c. and this to 50 c.c. Note the odor of this distillate. Test its inflammability in a warm, shallow dish. If it does not burn, distill again, and test the first 10 c.c. with a match. Perform the iodoform test for alcohol as follows: Make a 1.0 per cent solution of alcohol from the reagent bottle. Put 10 c.c. of this in a test tube, and 10 c.c. of the distillate obtained above in another test tube. Into each tube put 2 or 3 very small crystals of iodine, then 5 c.c. dilute sodium hydroxide. Warm gently. A yellow precipitate, or a yellow color, and a characteristic iodoform odor, or " hospital smell," indicate the presence of alcohol in the original solution. Write the following equations: (1) The action of the carbon dioxide in the limewater. If the precipitate that first formed in the cylinder redissolved later, see page 102 for the reaction. (2) The fermentation of dextrose into alcohol and carbon dioxide.

24. To Study Formaldehyde.— (a) Into some methyl (wood) alcohol in a test tube drop a red-hot coil of copper wire. After the flame has ceased note the odor of formaldehyde. The heating of the copper causes it to oxidize slightly, forming copper oxide; the oxygen of this oxide then combines with 2 hydrogens of the alcohol. Note the brightness of the copper after the experiment. The equation is:

$$CH_4O \;+\; O \;=\; CH_2O \;+\; H_2O$$
$$\text{methyl alcohol} \qquad\qquad \text{formaldehyde}$$
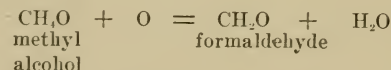
(b) Add about 5 g. of potassium permanganate to 10 c.c. of formalin in a beaker. Note the evolution of formaldehyde gas. This is the common way of generating the gas for fumigation. What is formalin?

25. To Prepare Some Esters.— (a) Mix a few c.c. each of ethyl alcohol and of acetic acid in a test tube. Add slowly 1 or 2 c.c. concentrated sulfuric acid, mix gently and warm. Note the fruity odor of *ethyl acetate*.

(b) Repeat, using ethyl alcohol and butyric acid. What fruit does this ester suggest?

(c) Repeat, using methyl alcohol and salicylic acid. What flavor is this ester?

ACIDS, ALKALIES AND SALTS

FROM time to time in the preceding chapters we have had occasion to mention several acids, such as sulfuric, nitric, hydrochloric, and various fruit acids. The term acid is familiar to most people; ask people who have never studied chemistry what an acid is, and a great many will say that an acid is something which tastes sour. And that is a fairly correct answer. For we may have the taste of substances described as bitter, sweet, biting, astringent; but the word sour conveys immediately a suggestion of green apples or cherries, vinegar, or sauerkraut, and these things we know are all acid-containing substances.

There are other properties of acids besides sourness, however; and as acids form a very important class of chemical substances, and as a study of the properties of acids leads us to a study of the other important compounds, we shall now consider acids in some detail.

Preparation of Acids.—A simple method of preparing an acid for study is to repeat the experiment in Chapter III, where a piece of phosphorus is burned under a bell jar standing in water. In this case it is easier to proceed as follows: a half inch of water is put into a wide-mouthed bottle provided with a cork. (A glass one-pint fruit jar will do.) A flat cork is floated in the water and a piece of phosphorus the size of a pea placed on the cork. The phosphorus is lighted, and the bottle closed. When the phosphorus has ceased burning the bottle and contents are

shaken until all the fumes of the burned phosphorus are
dissolved in the water. The water is then poured into a
beaker. This water now has a sour acid taste, suggesting
that the burned phosphorus had dissolved in the water to
form an acid. Although we have mentioned before the
chemical equations involved in this experiment, we repeat
them here:

$$\underset{\text{phosphorus}}{2\,P} \;+\; \underset{\text{oxygen}}{5\,O} \;=\; \underset{\text{phosphorus pentoxide}}{P_2O_5}$$

$$\underset{\text{phosphorus pentoxide}}{P_2O_5} \;+\; \underset{\text{water}}{3\,H_2O} \;=\; \underset{\text{phosphoric acid}}{2\,H_3\,PO_4}$$
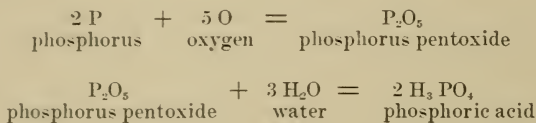
It is the phosphoric acid, H_3PO_4, that gives the water the
sour taste.

If, now, we put a piece of blue litmus paper into the
water, the paper will be turned pink. This is another test
for acidity; in fact, it is the commonest chemical test for
acids which we have. Litmus is the colored juice of
certain lichens; when strips of filter paper are soaked in it
and dried, we have our ordinary litmus test-paper.

The above experiment may be repeated, using sulfur;
an acid solution will again be obtained. It has already
been explained how nitric acid is produced in thunder-
storms (p. 20) by the oxidation of nitrogen and then
the solution of the nitrogen oxide in water. The common
name for carbon dioxide is carbonic acid or carbonic acid
gas; this is because solutions of the gas in water taste
somewhat sour, as in all of our soft drinks. The carbon
dioxide actually forms a compound with the water of
the formula H_2CO_3.

Examples of Acids.—We might name many other acids
and their methods of formation, but we have named
enough of them to be able to point out certain other charac-

teristics, especially if we assemble a number of common acids with their formulas in a list:

Phosphoric acid	H_3PO_4
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Carbonic acid	H_2CO_3
Hydrochloric acid	HCl
Acetic acid	$HC_2H_3O_2$
Tartaric acid	$H_2C_4H_4O_6$

Properties of Acids.—If we look at the formulas of the above acids, we will see that *they all contain hydrogen.* Some have one atom, some two, some three. These hydrogen atoms are called *acid hydrogens.* In the formulas of the last two acids the hydrogens are written in two groups; only the first ones, however, are to be considered as acid hydrogens, for reasons which will soon appear.

On looking again at the above formulas, we see that other elements making up the acid molecules are never metals, but are always non-metallic elements. Thus, in the above acids we see phosphorus, nitrogen, sulfur, carbon, and chlorine. Oxygen is sometimes, but not always, present. We can make a general statement to the effect that *every acid is a compound of a non-metallic element.*

Summary of Acids.—We have now mentioned several characteristics of acids, which serve to distinguish this group of compounds from all other chemical substances. A summary of these characteristics is as follows:

1. Acids are sour to the taste.

2. Acids turn blue litmus paper red.

3. The molecule of every acid contains at least one hydrogen atom.

4. Every acid is a compound of a non-metallic element.

Preparation of Alkalies.—If we now consider some of the characteristic features of compounds of metals, we find something entirely different from the above. If we take a piece of sodium, which is a metal, although not

used as such, like iron, and place it in a little water in a beaker, there will be a violent reaction, with the final disappearance of the sodium. If the pink litmus paper obtained in the acid experiment is put into the water, it will turn blue. If the water is tasted it will be found to have not a sour taste, but a brackish, caustic, puckering taste. The sodium has evidently combined with the water to form a compound entirely different from an acid. *This compound, which turns pink litmus paper blue, is called an alkali, or a base.*

The reaction involved with the sodium is as follows:

$$Na \quad + \quad H_2O \quad = \quad NaOH \quad + \quad H$$
sodium   water   sodium hydrate   hydrogen
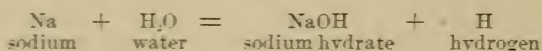
The hydrogen escapes as a gas, and leaves the *sodium hydrate* dissolved in the water. This sodium hydroxide is the ordinary caustic alkali, or caustic soda. Another alkali can be made by dissolving a little quicklime in water:

$$CaO \quad + \quad H_2O \quad = \quad Ca(OH)_2$$
Quicklime (calcium oxide)   water   calcium hydrate

This solution of *calcium hydrate* is a mild alkali; it is the common limewater often taken for stomach trouble.

A list of some common alkalies and their formulas follows:

Sodium hydrate	$NaOH$
Potassium hydrate	KOH
Iron hydrate	$Fe(OH)_3$
Copper hydrate	$Cu(OH)_2$
Calcium hydrate	$Ca(OH)_2$
Ammonium hydrate	NH_4OH
Lead hydrate	$Pb(OH)_2$

Properties of Alkalies.—Let us examine this list as we did the list of acids. In the first place we see that the alkalies do not contain a hydrogen alone, as did the acids,

but they do contain a hydrogen and an oxygen together. This OH is the *hydrate* group. Thus every alkali is a hydrate. The term hydrate comes from the fact that such a compound is usually formed from water, as in the case of the sodium and the calcium hydrates above.

The next thing we notice is that each of the above alkalies is a compound of a metal and not a non-metal. There are, in the above list, sodium, potassium, iron, copper, calcium, and lead. The nitrogen of the ammonia is not a metal, to be sure; but ammonium hydrate turns litmus blue, and the NH_4 group acts like a metal chemically in a great many ways. The above list might have been extended to include almost all of the metallic elements. It must not be thought that all of them form hydrates that are soluble in water, like those of sodium and calcium. But the general statement can be made that *the alkalies are compounds of the metals with the hydrate group.*

Summary of Alkalies.—We can now summarize the characteristics of alkalies as we did those of acids:

1. Alkalies do not taste sour, but cause a brackish, puckering sensation in the mouth.

2. Alkalies turn litmus paper blue.

3. The molecule of every alkali contains at least one hydroxyl, or OH, group.

4. Every alkali is a compound of a metallic element.

Formation of Salts.—If a piece of litmus paper is dropped into a little dilute acid in a beaker and some dilute alkali is added, a little at a time, with constant stirring, the litmus paper will remain red until a certain amount of alkali has been added, when it will turn blue. Evidently the acid has been destroyed, so that alkali now exists in the beaker and exerts its effect on the litmus instead of the acid. If the alkali is added very carefully, a point can be reached where the litmus is neither red nor

blue, but a tint intermediate between them. *This point, where a solution is neither acid nor alkaline, is called the neutral point.* That is, the alkali has neutralized the acid; or conversely, the acid has neutralized the alkali.

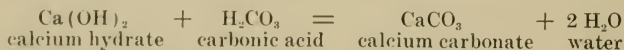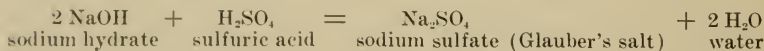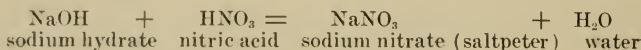Let us illustrate this by means of sodium hydrate and hydrochloric acid, and use a chemical equation:

$$\underset{\text{sodium hydrate}}{\text{NaOH}} \quad + \quad \underset{\text{hydrogen chloride}}{\text{HCl}} \quad = \quad \underset{\text{sodium chloride}}{\text{NaCl}} \quad + \quad \underset{\text{water}}{\text{H}_2\text{O}}$$
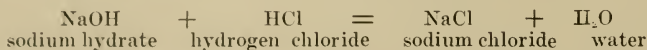
The hydrogen of the acid combines with the hydrate group of the alkali to form water, which leaves the non-metallic element, chlorine, of the acid, to react with the metallic element, sodium, of the alkali, to form the sodium chloride. This latter compound will be recognized as common salt. In fact, it is *a* salt; that is, it is one of a large class of substances called salts. All salts are combinations of a metal with a non-metal, hence a great many different salts are possible. Epsom salt, saltpeter, Glauber's salt, are common examples of salts other than common table salt.

Neutralization.—The above equation is typical of the chemical changes that always take place when an acid and an alkali act on each other. That is, *when an acid and an alkali neutralize each other, a salt and water are always formed.* A few other examples of neutralization will be given:

$$\underset{\text{sodium hydrate}}{\text{NaOH}} \quad + \quad \underset{\text{nitric acid}}{\text{HNO}_3} = \quad \underset{\text{sodium nitrate (saltpeter)}}{\text{NaNO}_3} \quad + \quad \underset{\text{water}}{\text{H}_2\text{O}}$$

$$\underset{\text{sodium hydrate}}{2\,\text{NaOH}} \quad + \quad \underset{\text{sulfuric acid}}{\text{H}_2\text{SO}_4} \quad = \quad \underset{\text{sodium sulfate (Glauber's salt)}}{\text{Na}_2\text{SO}_4} \quad + \quad \underset{\text{water}}{2\,\text{H}_2\text{O}}$$

$$\underset{\text{calcium hydrate}}{\text{Ca(OH)}_2} \quad + \quad \underset{\text{carbonic acid}}{\text{H}_2\text{CO}_3} \quad = \quad \underset{\text{calcium carbonate}}{\text{CaCO}_3} \quad + \quad \underset{\text{water}}{2\,\text{H}_2\text{O}}$$

It will be noticed that one acid hydrogen always combines with one hydrate group to form one molecule of water. If, as in the second equation, the acid happens

to have two hydrogens, and the alkali only one hydrate group, two molecules of the alkali have to be used to satisfy the acid.

In looking at the formation of a salt carefully, it will be seen that it actually consists of a metal replacing the hydrogen of the acid. Thus, some of the salts of hydrogen nitrate (nitric acid) are sodium nitrate, copper nitrate, lead nitrate, silver nitrate. Hydrogen sulfate is sulfuric acid; some of its salts are tin sulfate, nickel sulfate, potassium sulfate.

Neutralization is a very common and important chemical reaction. When a soil becomes acid, it is neutralized with lime. When a fruit or food is too sour, it is neutralized by means of baking soda, a mild alkali. If our stomach becomes too acid, we take limewater to partially neutralize the acidity. If we spill acid on our clothes or body, we are always told to put soda or ammonia on it to neutralize it. The buttermaker can measure the amount of acid formed in sour cream by finding out how much alkali is required to neutralize it.

Sugar is not an alkali. Therefore, when it is used to sweeten sour fruits, it does not neutralize their acidity; it simply masks their sourness, so that we taste the sugar instead of the acids.

Radicles.—In discussing acids we found that an acid consists essentially of a non-metal combined with hydrogen. The non-metal may exist alone with the hydrogen, as in HCl, where chlorine is the non-metal, or the non-metal may be linked with some oxygen atoms, as in HNO_3; or it may be linked with a still larger group, as in $H_2C_4H_4O_6$ (tartaric acid), where the four carbon atoms are the non-metal and are a part of the whole group $C_4H_4O_6$, only the first two hydrogens being acid hydrogens and replaceable by a metal. Thus some of the salts of tartaric acid are $CaC_4H_4O_6$, $K_2C_4H_4O_6$, $CuC_4H_4O_6$. *These*
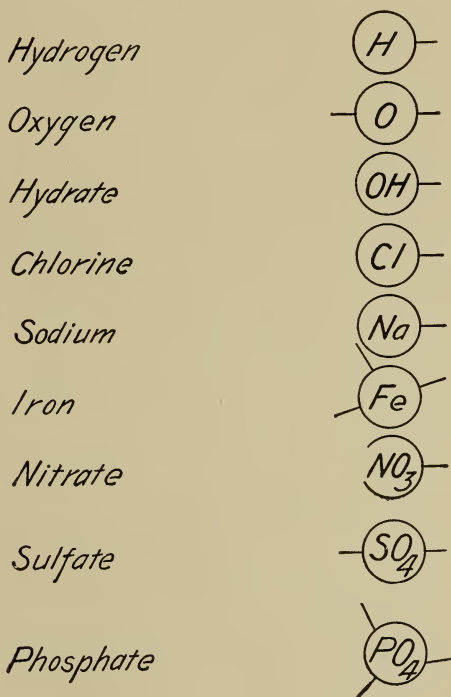
groups are called the acid radicles. Thus the nitrate radicle is $-NO_3$, the phosphate radicle is $-PO_4$, the acetate radicle is $-C_2H_3O_2$, and the sulfate radicle is $-SO_4$.

Instead of defining a salt as the combination of a metal and a non-metal, we can now more accurately say that *a salt is a compound consisting of a metal combined with an acid radicle.* Another way of stating it is that *a salt is an acid with the hydrogens replaced by a metal.* We speak of the salts of sulfuric acid, meaning the compounds which are formed when various metals replace the hydrogens of sulfuric acid. Or, we speak of the salts of copper, meaning the compounds which arise when copper replaces the hydrogen of various acids.

Metals and Non-metals.—Up to now we have not attempted to distinguish between the chemical elements that are metals and those that are not metals. The chemical elements can be roughly divided into these two classes. The metals are characterized in general by being solids, by having a metallic lustre, by conducting heat and electricity readily, and by being capable of being drawn into wires. All the metals do not conform to all of these characteristics, to be sure, but in general these are true. From what we have just seen in connection with acids and alkalies, we can give a further chemical distinction between the metals and non-metals, and that is that *non-metals form acid radicles, and the metals replace the hydrogen of acid.* Thus, the chemical behavior decides whether an element is to be considered a metal or a non-metal. Sodium does not look any more like a metal than does phosphorus; but as it forms a strong alkali, and as phosphorus forms a strong acid, there is no mistaking the class to which each belongs.

Valence.—The student by now may have noticed certain irregularities in the proportions in which the various

elements and radicles are combined. We have NaOH, Ca(OH)$_2$ and Fe(OH)$_3$; we have HNO$_3$, H$_2$SO$_4$, and H$_3$PO$_4$; we have CO$_2$, H$_2$O, and SO$_3$; we have NaCl, CaCl$_2$, FeCl$_3$. That is, two hydroxyls or two chlorines combine with one calcium, whereas it requires three of them to combine with one iron, and only one for a sodium, and so on. Different elements and radicles thus have different combining power. Suppose we represent each of these by a circle, with an arm on it for each combining power. We would have something like these:

Hydrogen   (H)—

Oxygen   —(O)—

Hydrate   (OH)—

Chlorine   (Cl)—

Sodium   (Na)—

Iron   (Fe)

Nitrate   (NO$_3$)—

Sulfate   —(SO$_4$)—

Phosphate   (PO$_4$)

When these atoms and radicles combine, let us imagine them joining hands. A hand of one never grasps more

than one hand of another, and all hands must be occupied when the compound is complete. Some of the above might combine as follows:

water   hydrochloric acid   calcium chloride
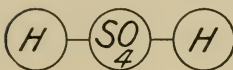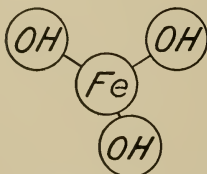
sulfuric acid   iron hydrate   iron phosphate
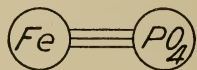
It must not be imagined that these atoms and radicles really have arms thrust out ready to unite with others; we have never yet been able to see an atom or a radicle, and do not know what the mechanism of chemical combinations is. But the above method illustrates the behavior of the elements, and it is useful for that reason. *This combining power of elements and radicles has been given the name of valence.* The hydrogen atom has a valence of one, the sulfate radicle two, and so on.

TABLE V

Valences of the Commoner Elements and Radicles

Valence of I	Valence of II	Valence of III	Valence of IV
H	Mg	Al	
NH$_4$	Ca	Fe	
Na	Fe		
K	Cu		
Ag	Zn		
Cl	S	PO$_4$	Si
I	SO$_4$		
NO$_3$	CO$_3$		
OH	O	N	C

If calcium has a valence of two, and chlorine of one, it will take two of the latter to satisfy one calcium. Therefore, the formula is $CaCl_2$, and not $CaCl$ or Ca_2Cl.

It is very important that the student memorize the valence of the more common elements and radicles. These are given in Table V. Only the simpler ones are given; some elements have two valences, and some have a valence as high as eight; but it would be useless to go into that here.

The next table summarizes the characteristics of acids and alkalies.

TABLE VI.

Summary of Characteristics of Acids and Alkalies.

Acids	Alkalies
1. Turn blue litmus paper red.	1. Turn red litmus paper blue.
2. Taste sour.	2. Taste soapy and slippery.
3. Contain H combined with a non-metallic element or radicle.	3. Contain OH combined with a metallic element or radicle.
4. Combine with alkalies to form water and salts.	4. Combined with acids to form water and salts.

QUESTIONS

1. Write equations showing the reaction involved when each of the first five acids in the list on page 85 neutralizes each of the first five alkalies on page 86. Name the salt formed in each case. Pay close attention to the valence of each element and radicle.
2. How could you neutralize some vinegar in a beaker? How would you know when it is neutral?
3. How would you prepare calcium phosphate?
4. What is the difference between a metallic and a non-metallic element?
5. What is an acid hydrogen?
6. Explain why a fruit no longer tastes so sour after sugar is added to it.

LABORATORY EXPERIMENTS

26. **To Study the Properties of Acids.**—Prepare *very dilute* solutions of hydrochloric, nitric, acetic and tartaric acids. Test each one as to taste, effect on sodium bicarbonate, effect on litmus paper and on solutions of other indicators such as cochineal and phenolphthalein. To what element in the acids are these characteristics due? What other property of acids was shown in experiment 12?

27. **To Study the Properties of Alkalies.**—(a) Prepare *very dilute* solutions of sodium hydroxide, calcium hydroxide (limewater), and ammonium

hydroxide. Perform the same tests on them as were performed on acids in experiment 26. To what are the characteristics due in this case?

(b) Take some juice of apple, orange, or lemon, and divide it into two parts. Carefully add some very dilute sodium hydroxide to one portion until it barely turns litmus paper pink. Compare the taste of the two portions of juice. What has become of the acids in the treated portion?

28. To Show the Colors of Indicators.—An indicator is a substance which has a different color in acid from what it has in alkali. Record the color of litmus, methyl orange, cochineal, phenolphthalein, and any other indicators that are available, in acid and in alkaline solutions. Dip bruised leaves of purple cabbage in weak acid and in weak alkali and record the results. Hold various colored flowers over an open bottle of ammonia and an open bottle of hydrochloric acid. What would you say as to the reaction (acid or alkaline) of the sap of these flowers?

29. The Preparation of Common Salt.—Dilute 50 c.c. of concentrated hydrochloric acid with an equal volume of water. Then carefully add sodium hydroxide solution (15 or 20 per cent) with constant stirring with a glass rod until the solution just turns blue litmus paper pink. Evaporate this solution down to dryness in a porcelain evaporating dish. Taste the residue. What is it? Write the equation showing its formation. What is the process of destroying an acid by an alkali called? What would have been formed in the case of each of the following acids acting on each of the alkalies: HNO_3 and H_2SO_4 on $NaOH$, KOH, $Ca(OH)_2$, NH_4OH? Write equations.

CHAPTER VIII

THE LIGHT METALS

Two Groups of Elements.—In the second chapter we learned that there are some eighty chemical elements known and that of these eighty only thirty are at all common. In this brief course, covering the chemistry of those elements which make up the substances in the little world in which we have our daily lives, we cannot study more than twenty of the elements. With these twenty, however, we ought to be very familiar. They make up the air we breathe, the water we drink, our food, clothing, building materials, the soil, the rocks, our drugs and medicines, our coins, paper, ink, toilet articles; in short, everything that we ordinarily have to use. It is therefore important that we now take up a systematic study of these few chemical elements. About each one we want to know where and how it is found in nature, how we utilize it, and what its important compounds are. We naturally divide this group of elements into metals and non-metals, and shall consider each class separately.

Simply for convenience, we shall divide the metals into groups, as follows:

I. *Light metals.*

1. Alkali metals. These include sodium and potassium and several rare elements. They are called alkali metals because their hydrates are our most powerful alkalies. We mentioned NaOH and KOH in the last chapter.

2. Alkaline earth metals. This group includes calcium and magnesium, and the less important metals strontium and barium. They receive the name of alkaline earths

from the facts that they form fairly strong alkalies, and that they are very abundant in the earth.

3. Aluminum.

II. *Heavy metals.*

4. Iron and others. This is not a definite chemical group; under this heading we will discuss iron, lead, copper, tin, zinc, and mercury.

5. The precious metals, silver, gold, and platinum. These metals are not only in limited amounts in the earth, but they are also very useful; hence they are called the precious metals.

1. Alkali Metals.—Sodium.—Sodium is a very abundant metal, constituting over two per cent of the earth's crust. It never occurs as the free element, for it has a powerful chemical attraction for many other elements. We saw in the last chapter that it combines violently with water. Common salt, NaCl, is its most abundant compound; sea-water is over three per cent sodium chloride, Great Salt Lake water contains from fifteen to twenty per cent of it; there are many other large salt lakes and marshes in the world, and there are enormous beds of the dry salt in many parts of the world, particularly in Michigan and New York in this country. This salt is used as the source of most of the other compounds of sodium.

Sodium Hydrate.—Although chemically sodium is a metal, the element itself cannot be used like other metals. It is soft, like putty, oxidizes to a powder readily in the air, and combines with water to form sodium hydrate. This sodium hydrate or *caustic soda* is used in very large quantities for making soap, and is used in small amounts in liquid soaps, which gives the latter their roughening effect on the skin.

Salts of Sodium.—As we said in the last chapter, the characteristic of all metals is that they form salts by com-

bining with acids. The salts of sodium are many and useful. With hydrochloric acid it forms the *sodium chloride* mentioned above. This is used for flavoring foods, for preserving meats and fish, for making a freezing mixture with ice, for glazing pottery. With nitric acid sodium forms *sodium nitrate,* $NaNO_3$, or Chile saltpeter. This is found in large beds in Chile; it is used mostly as a fertilizer because of its nitrogen content, although a great deal is mined for use in making nitric acid. *Sodium sulfate,* Na_2SO_4, is called Glauber's salt; it is commonly used as a veterinary laxative, in making glass, and in making sodium carbonate. *Sodium carbonate,* Na_2CO_3, commonly called washing soda or sal-soda, is the main constituent in washing powders and in water softeners. It is also used in large amounts in making glass and soap. The "black alkali" of some soils is due to sodium carbonate. *Sodium bi-carbonate,* $NaHCO_3$, is baking soda. It is much used in cooking, and is always one of the constituents of baking powders. *Sodium hyposulfite,* $Na_2S_2O_3$, is the ordinary "hypo" fixative of photography. Borax is *sodium borate,* used in softening water and for other laundry purposes. *Sodium silicate* is the "water-glass" with which we preserve eggs. All ordinary *hard soaps* are salts of sodium with certain acids obtained from fats. *Sodium compounds are found in the bodies of all plants and animals; they are not essential to the plants, however.*

Potassium.—The compounds of potassium are mostly duplicates of those of sodium, since the two metals are very similar in their chemical properties. Potassium compounds, however, are not nearly so easily obtained as those of sodium, as most of the potassium in the earth is combined in the granite and feldspar rocks. All soils contain small amounts of potassium. The one great source of potassium salts is the Stassfurt mines in Germany; and since the whole world was dependent on these

7

mines for its potash, the World War created potash starvation in many countries. The United States had to discover some other sources of supply, and succeeded in obtaining fairly large amounts from sea-weeds, from blast-furnace residues, and from certain saline lakes in California. Sea-weeds are high in both potassium and iodine, a substance much used in disinfecting wounds. The weeds are gathered, dried, and burned, and the potassium and iodine compounds extracted from the ash with water.

Potassium Hydrate.—*Metallic potassium* is very similar in its properties to sodium. With water it forms *potassium hydrate,* KOH, the powerful caustic alkali mentioned in the last chapter. It is used in dehorning cattle to some extent, also in making soft soap.

Salts of Potassium.—*Potassium chloride* is found in small amounts in sea-water, and in larger amounts in the Stassfurt beds. Its commercial name is muriate of potash, and is largely used as a fertilizer. *Potassium nitrate,* KNO_3, is ordinary saltpeter, used in corning beef, as a fertilizer, and in making black gunpowder. In the latter case it serves as an oxidizing agent to oxidize the other two constituents of the powder, charcoal and sulfur. *Potassium sulfate,* K_2SO_4, is a more common fertilizing material than the chloride; it is also used in making hard glass. *Potassium carbonate,* K_2CO_3, is the lye from wood ashes, used in making soft soap. The latter is the potassium salt of the acids found in fats. Fresh hard wood ashes contain from six to eight per cent of K_2CO_3, and hence are valuable fertilizing material. Potassium cyanide, KCN, as well as sodium cyanide, NaCN, are used in generating hydrocyanic acid gas for fumigation (Fig. 27). Thus:

$$2KCN \quad + \quad H_2SO_4 \quad = \quad 2HCN \quad + \quad K_2SO_4$$
potassium cyanide   sulfuric acid   hydrocyanic acid   potassium sulfate
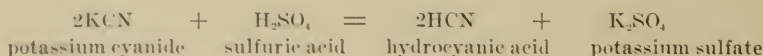
Potassium compounds are found in all plant and animal bodies, and *are essential to the life of all organisms.*

2. **The Alkaline Earth Metals.—Calcium.—***Calcium* and *magnesium* occur mostly as their carbonates in nature. *Calcium carbonate,* $CaCO_3$, is limestone. We have discussed in a previous chapter the origin of the immense
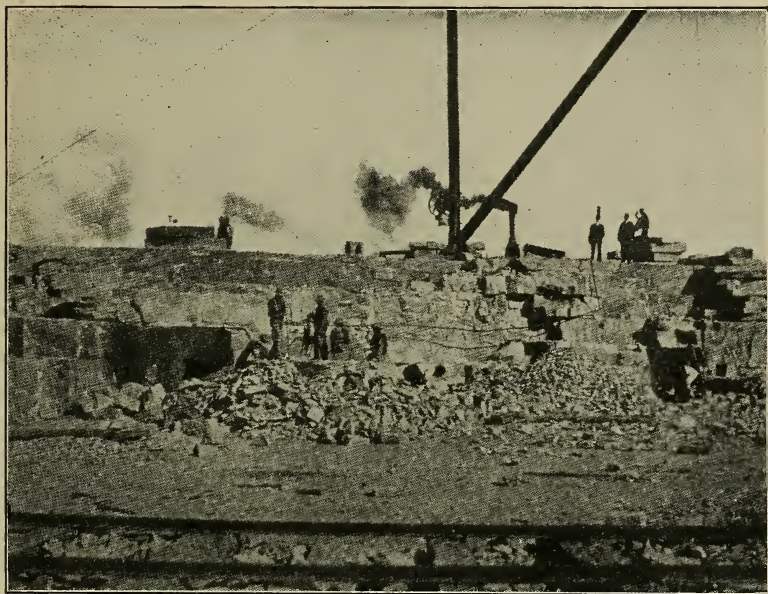
Fig. 28.—A limestone quarry. Such beds of calcium carbonate are the result of accumulation of the skeletons of minute sea-animals for thousands of years. That these land areas were once the bottom of the ocean is proved by the fact that fossils of sea-animals are frequently found in them. (From Bul. 230, Wis. Agri. Exp. Station.)

limestone, chalk, and marble beds (Fig. 28). Marble is crystallized limestone; the change was brought about by great heat and pressure in the earth. Chalk is another form of calcium carbonate found in nature. Both marble and limestone are valuable building materials.

Quicklime.—Limestone finds its greatest use, how-

ever, as the source of quicklime, and thus of our building mortar and our whitewash. Lime plaster is such a familiar substance that its chemistry should be understood from beginning to end. When the limestone is "burned" or roasted, in kilns or piles, as in figure 29, to about

Fig. 29.—A home-made kiln for burning oyster shells to quicklime. Similar improvised outfits can be used for burning other forms of calcium carbonate, such as clam shells from button factories, limestone, marl and marble. (From U. S. Dept. of Agriculture, Farmers' Bul. 921.)

$1100°$ F., carbon dioxide is driven off and calcium oxide is left behind:

$$CaCO_3 = CaO + CO_2$$
$$\text{calcium carbonate} \quad \text{calcium oxide} \quad \text{carbon dioxide}$$

The carbon dioxide escapes from the vent at the top. The calcium oxide is *quicklime*. One hundred pounds of

good limestone make about fifty pounds of quicklime. The latter is in hard, porous lumps. It can be air-slaked or water-slaked. In the former it re-absorbs from the air the carbon dioxide it lost:

$$CaO \ + \ CO_2 \ = \ CaCO_3$$
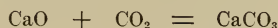
Hence air-slaked lime is calcium carbonate, or the same chemically as limestone itself. When quicklime is slaked in water it unites with it as follows:
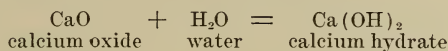
$$\underset{\text{calcium oxide}}{CaO} \ + \ \underset{\text{water}}{H_2O} \ = \ \underset{\text{calcium hydrate}}{Ca(OH)_2}$$

If the proper amount of water has been added (100 pounds of lime to 30 pounds of water) the lime will crumble into a dry powder. This is the hydrated lime of commerce. This will dissolve to a slight extent in water to form a clear solution called *limewater*. This is a mild alkali used medicinally. If larger amounts of the hydrated lime are made into a thin paste with water it is called *milk of lime*.

Lime Plaster.—It is this milk of lime which is mixed with sand and other materials for mortar. The wet mortar is spread on walls or between stones or bricks and allowed to ''set.'' This wet mortar chemically is a mixture of sand, water and $Ca(OH)_2$. As the water evaporates it leaves the mortar porous; carbon dioxide from the air can then seep in and react with the calcium hydrate as follows:

$$Ca(OH)_2 \ + \ CO_2 \ = CaCO_3 \ + \ H_2O$$

Thus more water is produced; and at the same time calcium carbonate, the same substance as the original limestone, is gradually formed around the particles of sand. By the time the water has all evaporated and all the calcium hydrate has become carbonate, the mortar is

"set." It consists now of a mixture of calcium carbonate and sand, together with any coloring matter, hair, or other substance that was added to it. We now see that the calcium has "gone around in a circle," for it started out as the carbonate in the original limestone, then went to the oxide, then to the hydrate, and ended up as the carbonate again. Looking back at the source of the limestone beds, we can consider that the calcium carbonate that once formed the skeleton framework of sea animals now forms part of the framework of our dwellings.

Origin of Hard Water.—There is another very important relation between calcium carbonate and carbon dioxide. The carbonate is practically insoluble in water; therefore water alone in seeping through limestone beds would not dissolve much of it. But as water passes through the soil, it absorbs considerable carbon dioxide which is formed in the soil by the decay of organic matter. This water charged with carbon dioxide attacks the limestone in the following way:

$$CaCO_3 \quad + \quad CO_2 \quad + \quad H_2O = \quad Ca(HCO_3)_2$$
calcium carbonate   carbon dioxide   water   calcium bicarbonate
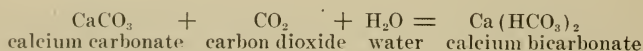
This calcium bicarbonate dissolves readily in the water. The latter is now a "hard water." It forms a curd with soap, and leaves a residue in the bottom of tea-kettles and boilers (Fig. 30). When water containing calcium bicarbonate is boiled, the reverse of the above equation takes place:

$$Ca(HCO_3)_2 \quad = \quad CaCO_3 \quad + \quad CO_2 \quad + \quad H_2O$$
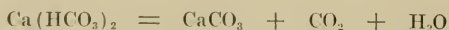
The carbon dioxide is expelled, leaving the original calcium carbonate in the water. As it is insoluble, it settles out as a sediment; and as the water no longer contains any calcium salt in solution it is now a "soft water." *Such hardness of water, which can be removed by boiling,*

is called "temporary hardness." If the calcium salt in the water is calcium sulfate or chloride, it cannot be removed by boiling, and hence is "permanently hard."

Formation of Caves.—The Mammoth Cave in Kentucky, as well as other caves in limestone regions, was made by carbonated water dissolving out big pockets in the limestone, according to the equation given.

Plaster of Paris.—Besides the plaster involving calcium carbonate, described above, there is another plaster mate-
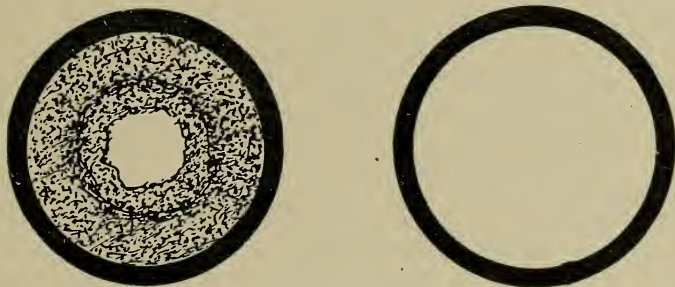
Fig. 30.—Sections of steam or hot water pipes with deposit of lime from the use of hard water.

rial which is a calcium compound. Plaster of Paris has the formula $CaSO_4\frac{1}{2}H_2O$; that is, the molecule of calcium sulfate has some water combined with it. When it is mixed with water to form a paste, it gradually combines with this water chemically to form a different combination of calcium sulfate and water. Thus:

$$CaSO_4\tfrac{1}{2}H_2O \ + \ 1\tfrac{1}{2}\,H_2O \ = \ CaSO_4\,2H_2O$$
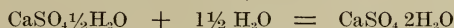
The latter compound is the set plaster. It is the same as our gypsum, or land plaster, and it is also the same as the alabaster of Biblical mention.

Portland cement is another calcium containing substance which sets on treating with water. In this case com-

plex silicates and aluminates are involved, into the chemistry of which we cannot go.

Calcium silicate, together with *sodium silicate,* are the main constituents of our ordinary window and bottle glass. In making it, finely ground sand, calcium carbonate and sodium carbonate are mixed together and fused in fireclay pots. The melted mass can then be poured into moulds or blown into any desired shape. Special kinds of glass contain potassium, borax, lead, cobalt, or other substances.

Calcium Phosphate.—The bones of animals consist almost entirely of *calcium phosphate,* $Ca_3(PO_4)_2$. This is also found in great beds in the Southern States, and is there called rock phosphate. It is much used as a phosphate fertilizer. *Calcium is also found in all plants and is a necessary element in their growth.*

Bleaching Powder.—We mentioned under the subject of purification of water another important compound of calcium, the *hypochlorite,* $CaOCl_2$; great quantities of it are used as a bleaching agent in paper-making, cotton and linen mills, and straw-hat factories.

Calcium Carbide and Sulfide.—In the chapter on carbon we discussed the manufacture and use of calcium carbide, CaC_2, and its reaction with water to form acetylene. Another interesting compound of calcium is the sulfide, CaS. After exposure to sunlight it will glow in the dark, and is therefore used in luminous paint on watch and compass faces and on the plates around keyholes.

Magnesium.—The next metal of the alkaline earth group, *magnesium,* has great interest to us as the element as well as in compounds. Metallic magnesium burns readily in the air with a brilliant white light. When it is finely powdered and mixed with potassium chlorate, a powerful oxidizing chemical, it burns in a fraction of a

second with a blinding flash of light. This mixture is the *flashlight powder* used in photography. Magnesium is also used in fireworks.

Occurrence of Magnesium.—In nature the *carbonate* of magnesium is called magnesite. A very common way in which magnesium carbonate occurs is in *dolomite,* which is a mineral composed half and half of calcium carbonate and magnesium carbonate, $CaCO_3.MgCO_3$. Other well-known minerals are *magnesium silicates,* as talc, soap-stone, meerschaum, and asbestos. Some magnesium salts are found in sea-water, having been dissolved from the soil. *It is a necessary element in both plant and animal life;* in fact, in plants it is a part of the chlorophyll, the green coloring matter, which is the substance that causes the union of carbon dioxide and water to form sugar.

Magnesium oxide, MgO, is used as a mild alkali in medicine under the names "magnesia" and "calcined magnesia."

Magnesium sulfate, $MgSO_4$, is commonly called Epsom salt, from the fact that it was first isolated from a spring at Epsom, England. It is used as a purgative. It occurs very commonly in water, causing permanent hardness. Often it occurs to such an extent in waters that they are used directly as laxatives. Many waters in the Western States prove to be so strong with magnesium sulfate that they cannot be used for drinking purposes. They are often called "alkali" waters, although chemically there is nothing alkaline about magnesium sulfate.

3. Aluminum.—Although aluminum minerals are exceedingly abundant in the earth, more abundant, in fact, than those of any other metal, the metal itself has come into daily use only during the last fifteen years, since an electrical process of manufacture has greatly reduced its cost.

Aluminum is a constituent of all clays, granites, micas, feldspars, and slates. The metal cannot be prepared from them, however; only two minerals, of less abundant occurrence than the above, can be used in its manufacture. The metal is silvery white, rather soft, very malleable and duc-

Fig. 31.—Scene in a clay mine. The clay is removed in a way similar to coal—by drilling, blasting, picking, and carrying out of the mine on cars. The above cut is from a Pennsylvania firebrick clay mine.

tile; it takes a fairly good polish, but the polish does not remain, as aluminum is slowly oxidized in the air with the formation of a grayish coating. When heated, aluminum unites very readily with oxygen; in fact, it will take the oxygen away from iron oxide, with the liberation of a great amount of heat. This property is utilized in

thermite welding. Thermite, which is a mixture of powdered iron oxide and aluminum, is placed around the two iron parts to be welded, such as the ends of rails, and the thermite ignited. The reaction proceeds with such violence that the ends of the rails are melted and fuse together. The reaction is as follows:

$$Fe_2O_3 \quad + \quad 2\,Al \quad = \quad Al_2O_3 \quad + \quad 2\,Fe$$

iron oxide   aluminum   aluminum oxide   iron
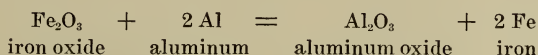
Aluminum is used in enormous quantities in making cooking utensils, cables for electric wires, and the frames of cameras.

Alum.—The *alums* are the most important salts of aluminum. There are many different alums, and they serve varied purposes, among which we might mention the dye industries, where alum helps to fix the dye in the cloth; water purification, where alum unites with the coloring matter and organic impurities and carries them down in a sludge; and in medicine as an astringent to check the flow of blood from wounds.

Clay.—The pottery and brick industries use compounds of aluminum found in all parts of the world, which go by the general name of *clays* (Figs. 31, 32 and 33). They are essentially aluminum silicates with some water combined with them to form part of their molecules. When wet, clays can be moulded into any desired shape; and when baked in ovens, they become hard and strong, and will not again soften with water. What the exact chemical change is that takes place during the baking is not known. The glaze on pottery has a composition similar to that of glass; it is formed by soda, lime, salt, and other substances reacting with the silica of the pottery during a second baking.

Fig. 32.—Moulding bricks preparatory to drying and baking. The raw clay is ground with water by machinery. It is then moulded by hand into the crude bricks at the left of the picture. The machine moulds the finished bricks and stamps on them the manufacturer's mark. They are then stacked on the floor to dry, and later are baked.

Fig. 33.—Kilns used for firing brick. They are brought to the kilns after drying for several days in well-ventilated sheds. The raw clay is essentially an aluminum silicate chemically combined with water. During the baking the water is lost and the clay is left in a condition in which it will not again take up the water. The fuel piped to these ovens is natural gas.

TABLE VII

Summary of the Properties of the Light Metals

| Metal	Where found and in what principal form	Uses of the metal and its compounds
Sodium	As NaCl in sea-water and in mines As Na_2SO_4 in alkali soils	The metal has no important industrial uses. NaCl—to flavor food; preservative; glazing pottery. Na_2SO_4—veterinary purgative. Na_2CO_3—washing soda; glass-making. $NaHCO_3$—baking soda. NaOH—soap-making. $NaNO_3$—Chile saltpeter; fertilizer; gunpowder. $Na_2S_2O_3$—hyposulfite in photography. Compounds of Na are essential to animal life but not to plant.
Potassium..	As KCl in sea-water and in Stassfurt mines in Germany In all granite rocks	The metal has no important industrial uses. KCl—fertilizer. K_2SO_4—fertilizer. K_2CO_3—in wood ashes; making soft soap. KOH—chemical industries. KNO_3—saltpeter; corning beef; explosives. KCN—fumigation for insects. Compounds of K are essential to both plant and animal life.
Calcium....	As $CaCO_3$ in limestone and marble As $CaSO_4.2H_2O$ in gypsum As $Ca_3(PO_4)_2$ in rock phosphate and apatite	The metal has no important industrial uses. CaO—quicklime, for lime mortar. $Ca(OH)_2$—limewater as medicine. $CaCO_3$—making quicklime; building stone; fertilizer. $CaSO_4.\frac{1}{2}H_2O$—plaster of Paris. $CaSO_4.2H_2O$—fertilizer. $Ca_3(PO_4)_2$—in bones and rock phosphate. CaC_2—carbide for making acetylene. Calcium compounds are essential to both plant and animal life.
Magnesium	As $MgCO_3$ in some limestones As $MgCl_2$ in the Stassfurt mines in Germany	Mg—used in flashlight powder. $Mg(OH)_2$—milk of magnesia; medicine. $MgCO_3.CaCO_3$—limestone for building. $MgSO_4$—Epsom salt; medicine. Magnesium compounds are essential to both plant and animal life.
Aluminum	As silicate in all clays	Al—used for cooking utensils. Alums—medicine; dyeing; water purification.

QUESTIONS

1. What are the differences between the metals and the non-metals?
2. What do you think are the three most important compounds of sodium? Why?
3. Why are potassium compounds more important than sodium compounds in the life of the world?
4. Why is KNO_3 a more valuable fertilizer than $NaNO_3$?
5. What is the most important use of $CaCO_3$?
6. Describe the chemistry of lime mortar from the limestone to the dry plaster.
7. What is the chemistry involved in the formation of a limestone cave?
8. What is marble? Chalk? Dolomite?
9. How can temporary hardness be removed from water? Use a chemical equation to show the reaction.
10. How is glass made?
11. Which is the most useful as the free element, Na, K, Ca, or Mg?
12. What do you think are the four most important compounds of calcium? Why?
13. What do you think are the three most important compounds of magnesium? Why?
14. Describe aluminum.
15. What are the most important uses of aluminum compounds?

LABORATORY EXPERIMENTS

30. **To Study the Properties of Aluminum.**— (a) Examine an aluminum utensil. Is aluminum lighter or heavier than iron, copper, or silver? For what is aluminum used besides cooking utensils?

(b) Obtain a piece of iron wire and one of aluminum, both of the same size and about .6 inches long. Holding one in each hand, insert the ends into a gas flame and note which conducts heat the faster. How is this property of advantage in cooking utensils?

(c) Heat some pieces of aluminum wire in 10 per cent sodium hydroxide solution in a test tube. Test the escaping gas with a match.

$$Al + NaOH + H_2O = Na\ AlO_2 + 3H$$
$$\text{sodium}\ \text{sodium}$$
$$\text{hydroxide}\ \text{aluminate}$$
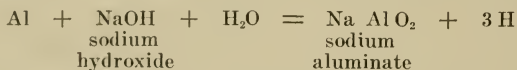
(d) Repeat (c), using 20 per cent hydrochloric acid.

$$Al + 3HCl = AlCl_3 + 3H$$

What are the two products formed?

31. **The Effect of Some Salts on Protein.**—For a protein solution, make a solution of 1 part egg white in 3 parts water. Pour about 5 c.c. of this solution into each of 7 test tubes. Test the solution with a few drops of the following salts: (1) sodium chloride, (2) potassium chloride, (3) calcium chloride, (4) aluminum sulfate (alum will do), (5) silver nitrate, (6) mercuric chloride ($HgCl_2$), (7) lead acetate. Record which metals form a precipitate with the egg white. Why is raw and not cooked egg white a

good antidote for mercury, lead, and silver poisoning? Explain the action of alum in stopping the bleeding of a cut.

32. To Prepare Plaster of Paris.—Heat some gypsum, $CaSO_4.2H_2O$, in a porcelain or iron dish in an oven at 110° to 115° C. for 15 minutes. Do not let the temperature go over 115°. Cool, mix with a small amount of water and let stand until it sets. Write equations showing the formation of the plaster and the setting of the plaster. What uses has plaster of Paris?

33. To Make Quicklime.—Heat some powdered limestone or marble in a porcelain crucible as hot as possible in the Bunsen flame for half an hour. Cool, then cover with water and note whether it slakes. If the slaking is not vigorous, it is because a sufficiently high temperature was not obtained for making quicklime. Test with litmus paper. An alkaline reaction shows the presence of some quicklime. Why? Write equations for the " burning " of the limestone and for the slaking in water.

34. To Study the Slaking of Lime.—Grind up about 50 g. of fresh quicklime into small pieces. Divide into two portions. Place one portion in a well-stoppered bottle and expose the other portion to the air in a shallow layer in an open dish. Two days later cover each portion with water in test tubes, warm slightly, and note which slakes the more readily. Explain the chemistry of air slaking and of water slaking, giving equations.

35. The Burning of Magnesium.—Hold a small piece of magnesium ribbon in the tongs and ignite in the flame. Write equation. For what is powdered magnesium used?

36. The Preparation of Iron Sulfate.—Place 25 g. of iron nails in a flask, add about 35 c.c. of dilute sulfuric acid (1 part acid to 5 parts water). Warm until no more gas is evolved. There should be some undissolved iron left. Filter the solution, and then evaporate to one-half its volume. Cool, and allow the iron sulfate to crystallize out. Remove and dry the crystals. Describe them. Write the equation. What are the uses of this compound? To what class of chemical compounds does it belong?

37. The Formation of Ink.—(a) Dissolve about 2 g. of the iron sulfate prepared in experiment 36 in 25 c.c. of water. Add a few c.c. of this solution to a weak solution of tannin. The colored compound is iron tannate, the basis of most fountain-pen inks.

(b) Make a strong extract of green tea by boiling a half teaspoonful of the leaves in a test tube of water. Test this extract with the solution of iron sulfate. Taste the solution of tannic acid prepared in (a), then explain the puckering effect of strong tea in the mouth.

38. The Effect of Carbon Dioxide on Calcium Carbonate.—Put a gram of powdered limestone into each of two test tubes, and add 10 c.c. of distilled water. Into one tube blow the breath through a piece of glass tubing for 5 minutes. Shake the other tube frequently. Filter each, and when the filtrates are clear, evaporate a 2 c.c. portion of each to dryness on watch crystals. Did the blowing of the breath increase the solubility of the calcium carbonate? How? Equation. How does this experiment illustrate conditions in the soil? How are caves formed? Boil the remainder of the filtrate containing the calcium bicarbonate in a test tube. Explain the origin of the precipitate. What has this experiment to do with hard water? Boil some hard water in a test tube and compare with the above.

CHAPTER IX

THE HEAVY METALS

The Iron Age.—Of all the dozens of metals that could be grouped under this heading, *iron* is preëminently the most useful and the most abundant. It occupies such a large place in our civilization that our age is called the ''Iron Age,'' in contrast to those of the earlier races of man, who lived in the Stone Age and the Bronze Age. Iron enters to a greater or less extent into all our structures, from skyscrapers to dwellings, and from bridges to watches; it forms the framework of railroad cars, steamships and automobiles; it forms the engines of all our means of locomotion (Fig. 34). It is the fourth most abundant element in the earth; it is present in the sun and in the other planets. The magnetism of the earth leads us to believe that the interior of the earth is mostly iron. Uncombined iron is found in only very small amounts in the earth, and this is mostly in meteors which have fallen to earth. The ores are very abundant and very widely distributed over the earth's surface. The usual ores used in smelting iron are the two oxides, hematite, Fe_2O_3, and magnetite, Fe_3O_4. The latter is magnetic; that is, it attracts other magnets and iron itself. A huge bed of it north of Canada constitutes the magnetic pole of the earth, to which all compass needles point.

Over 12,000 years ago the Greeks and Egyptians obtained small amounts of iron from its ores and forged it into weapons. The Romans, 2,000 years ago, made still more use of it. But iron has not been an abundant and cheap metal for much over 200 years, and it has not been used extensively for building purposes for over 50 years.

Smelting of Iron.—Since iron ores are compounds of iron and oxygen, the problem in obtaining the iron is to abstract the oxygen from it. We have seen that at high temperatures carbon has a strong affinity for oxygen. This property is utilized in smelting iron ores. A mixture of coke and iron ore is dumped into a large tower of fire-

FIG. 34.—A typical scene in the Iron Age in which we live. This is a large iron foundry housed in a building composed entirely of steel and glass.

brick called a blast furnace (Fig. 35). Since ore contains sand impurities, limestone is also added to the mixture as a flux. A blast of hot air is forced into the bottom of the furnace; the coke burns with great heat, combines with the oxygen of the ore, the iron which is freed is melted by the high temperature, sinks through the tower, and collects in a basin at the bottom. The limestone flux combines with the sand to form calcium silicate, which

8

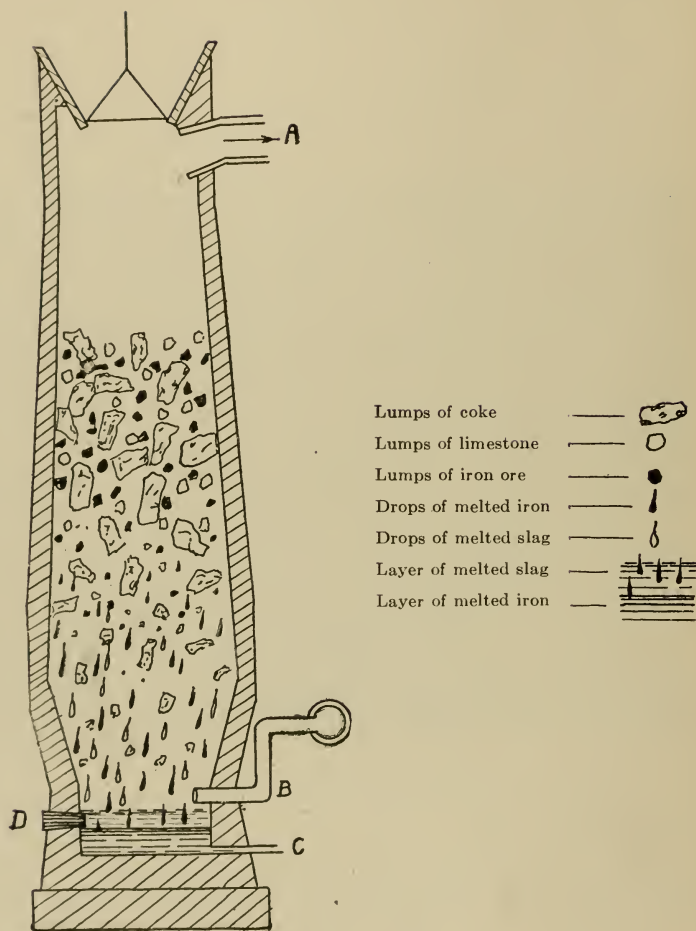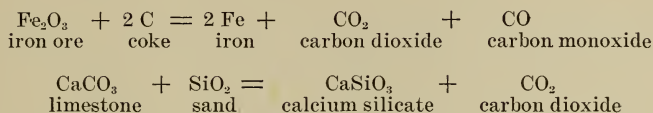|	
Lumps of coke	
Lumps of limestone	
Lumps of iron ore	
Drops of melted iron	
Drops of melted slag	
Layer of melted slag	
Layer of melted iron	

Fig. 35.—Diagram of a blast furnace, where iron is smelted from its ores. Ore, coke, and limestone are fed in at the top. A blast of highly heated air is blown in at *B*, which causes the combustion of the coke, the reduction of the ore, and the fusion of the limestone with the impurities of the ore to form a slag. The melted iron and slag drop to the bottom of the furnace; the iron is drawn off at *C*, the slag at *D*. The gases escaping at *A* contain a good deal of carbon monoxide; hence they are burned to heat the air blown in at *B*.

melts like glass, drops to the bottom of the tower and floats on top of the melted iron. The reactions involved are essentially as follows:

$$Fe_2O_3 + 2\,C = 2\,Fe + CO_2 + CO$$

iron ore   coke   iron   carbon dioxide   carbon monoxide

$$CaCO_3 + SiO_2 = CaSiO_3 + CO_2$$

limestone   sand   calcium silicate   carbon dioxide

It will be noticed that the gaseous products of the reaction are carbon dioxide and carbon monoxide, which pass out of the flue at the top. Since the CO is still capable of burning and furnishing considerable heat, the gases from the flue are brought down and burned to heat the air which passes into the furnace as a blast.

Kinds of Iron and Steel.—The blast furnace is operated continuously. Mixed ore, coke and limestone are added at the top from time to time, and the melted iron and melted slag are drawn off at the bottom through separate taps. The iron is run into moulds, and when it hardens it is the pig-iron of commerce. This pig-iron, or *cast-iron,* is used directly in castings, such as pipes, the heavy portions of machinery, pumps, and similar articles. It is very brittle and breaks readily. It is far from being pure iron, containing considerable carbon, sulfur, and phosphorus. When these are mostly removed, and the iron subjected to mechanical annealing, the product is *wrought-iron.* It is very tough, soft, and malleable. When the cast-iron is subjected to purification from carbon, sulfur and phosphorus, and then definite amounts of carbon added back to it, *steel* is the result. This is hard and has a crystalline structure. Its properties can be greatly modified by treatment with heat and by the amount of carbon added to it. Thus, as a result of special treatments, we can have very hard steel for tools and cutting edges, or we can have tough, elastic steel for springs and structural girders.

The properties of steel can be greatly modified by the addition of other metals; thus *nickel steel* is hard and elastic, and is used in the armor plate on warships; *manganese steel* is extremely hard, and is used for stone-crushers, safes, etc.; *tungsten steel* retains its magnetism very well, hence is used for permanent magnets; *vanadium steel* is tough and elastic and is much used in automobile parts.

Compounds of Iron.—The compounds of iron are numerous, but are relatively unimportant. The ore discussed above, Fe_2O_3, when finely ground, is used as a dark red paint pigment under the name of *Venetian red* or *red ochre*.

Iron sulfate, $FeSO_4$, also called green vitriol and copperas, is used for killing weeds, in the dye works, and in making writing ink. In the latter, the iron is combined with the tannic acid from nutgalls, forming *iron tannate,* which is bluish-black.

Laundry bluing is a complex compound of iron and potassium cyanide. A similar iron cyanide compound is used in making blueprint paper. *Iron citrate* is used as a medicine to supply iron to the blood. The *red corpuscles* of blood are composed of a complex compound of iron and protein. Its function is to carry oxygen from the lungs to the various tissues, where it is used for combustion. Iron therefore plays a very essential part in animal physiology. It is also always found in small amounts in plants, and is connected with the formation of the green coloring matter, chlorophyll.

Lead.—This metal is seldom found in the uncombined state in nature, but usually as the sulfide, PbS, commonly called galena. When this is roasted, oxygen combines with the sulfur and sets the lead free:

$$PbS \ + \ 2O \ = \ Pb \ + \ SO_2$$

Uses of Lead.—About a million tons of lead are annually produced in the world. Great quantities of it are used in making water pipes. Water containing dissolved air will dissolve lead slightly, and the soluble compound formed is somewhat poisonous. If, however, the water is at all hard, that is if it contains some calcium bicarbonate or sulfate, a coating of lead carbonate or lead sulfate is formed on the lead pipe which prevents any further dissolving of the metal. Hence it is perfectly safe to use lead pipes for drinking water, unless the water is soft. Other very important uses of lead are in storage batteries, babbitt metal, shot, solder, and type metal. The most important compound of lead is a carbonate, commonly called *white lead*. It has been the basis of paints for three hundred years, and the old Dutch process of making it, although slow, still produces a product with the best covering power. Briefly, the process consists in allowing carbon dioxide and the vapors of acetic acid to act on sheets of lead. The lead is suspended over an earthen jar of vinegar, which contains the acetic acid, and fermenting horse manure around the jar furnishes the carbon dioxide, and the heat necessary to vaporize the acid and the vinegar. The white lead ground with linseed oil forms the body of the paint; to this are added pigments to give paints of different colors. Some substitutes and adulterants of white lead are *lead sulfate* and *barium sulfate*, which form paints of inferior wearing qualities. Painter's sickness is a result of the skin absorbing some of the poisonous lead compounds. *Lead arsenate* is a poisonous compound used as a cheaper substitute for Paris green in killing potato beetles and other chewing insects.

Copper.—This metal is found in the uncombined state in considerable quantities; this fact, together with the fact that it is easily worked into various shapes by hammering, is the reason it was one of the earliest metals used by

man. Copper implements are found in the most ancient
of ruins; and the alloy of tin and copper known as *bronze*

FIG. 36.—Spraying poisonous compounds of copper and of arsenic on trees to kill insects
and fungous diseases. These compounds are applied to plants by means of sprayers varying
in size from a small hand pump to the powerful type shown in this picture. Acres of walnut,
pecan, orange, apple, and in fact all fruit and ever shade trees can be thoroughly and
cheaply sprayed in this way. Other types of power sprayers are designed to spray both the
under and upper leaf surfaces of potatoes, cabbages, and other low plants. (Courtesy of
U. S. Dept. of Agriculture, Bureau of Plant Industry.)

was for hundreds of years the most common metallic mate-
rial in use by man. Most of our copper to-day, however,
is smelted from its ores, which are chiefly the sulfides and

carbonates. Copper is a rather hard and tough metal; it can be hammered or rolled into sheets and drawn into wires readily, takes a high polish, and corrodes but slowly in the air. For these reasons it is used in vast quantities for roofing material, ornamental cornices, electric wires, covering the bottom of ships and for coins.

Copper Alloys.—It forms some of our most important alloys, among them being *brass,* which is copper and zinc; *German silver,* copper, zinc, nickel; *bronze,* copper and tin; our silver and gold coins are one-tenth copper, our nickels are three-fourths copper and one-fourth nickel, and our pennies contain a little tin and zinc besides the copper.

Copper sulfate is the best-known compound of copper; it is commonly called "blue stone" and "blue vitriol." Its formula is $CuSO_4$. It is used in preparing other copper compounds, in killing the weeds in ponds, in sterilizing swimming-tank water, and in making Bordeaux mixture, which is the common spray material used to protect plants from fungous diseases (Figs. 36 and 37).

Tin.—Tin is not found in the free state in nature, but usually as the oxide SnO_2, from which it is readily freed by roasting with coal, as in the smelting of iron. For this reason it could be produced even in ancient times and was used in connection with copper to produce bronzes. The famous tin mines of Cornwall, England, were first discovered and worked by the Romans during their occupation of England. Besides bronze, tin forms *solder* and *pewter* when alloyed with lead.

Tinware.—The greatest use of tin is in coating iron to prevent rust and the action of fruit acids. All our tinware, such as cans, kitchenware, and roofing, is made by dipping the iron core into melted tin. The tin itself will neither rust nor corrode; but when the thin coating of tin is marred, the iron beneath rusts very rapidly. Milk cans, separator parts, etc., have very heavy coatings of
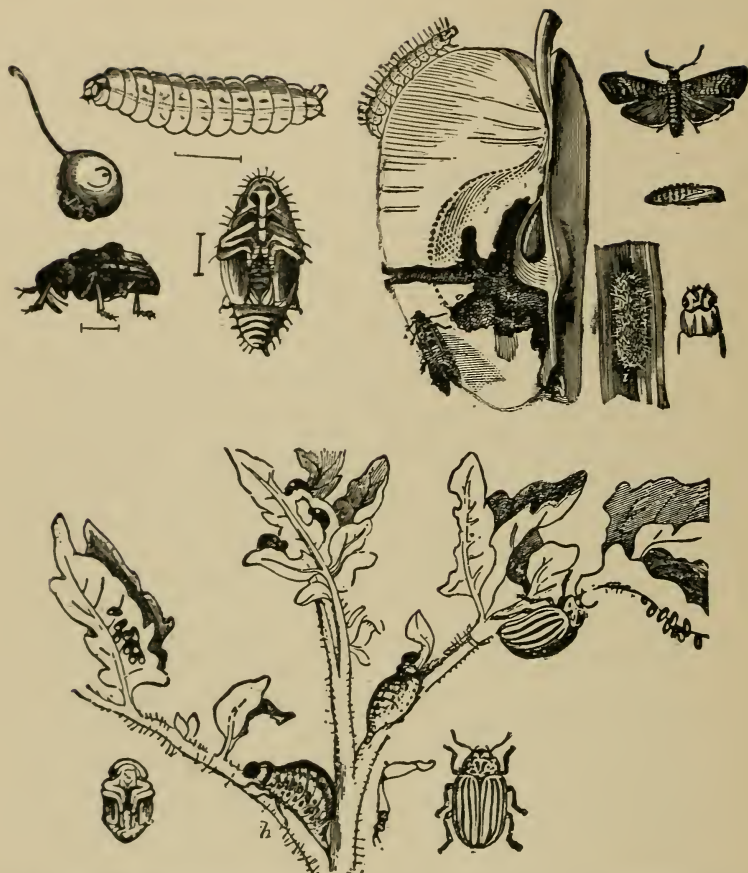
Fig. 37.—Spraying with poisonous compounds of arsenic and lead will largely control these insects. Left above, plum curculio with small fruit just attached. Right above, codling moth and its effect on apple. Below, potato beetle in different stages on the plant. (Smith's Insect Friends and Enemies.)

tin which prevent their rusting for a long time. The principal use for compounds of tin is in the weighting of silk.

Zinc.—This metal occurs usually as the carbonate and the sulfide. The ore is first roasted, then smelted with coke or coal. The metal is used in the sheet form for gutters, roofs, tanks, fancy gables, and other exposed parts, because it corrodes but very slightly in air. When a sheet of iron is dipped into melted zinc it receives a coating of the zinc. This material is then called *galvanized iron.* *Brass* is the principal alloy of zinc. *Zinc oxide,* ZnO, is used in huge quantities as a base for paint instead of white lead. As it is a mild poison, it is used in ointments as an antiseptic.

Mercury.—The only common metal that is a liquid at ordinary temperatures is mercury. It boils at a relatively low temperature, at about 500° F. It is used in barometers and thermometers. Since it freezes to a solid at 39° F. below zero, alcohol instead of mercury must be used in thermometers in cold climates. One of the most interesting properties of mercury is that of amalgamation with other metals. This is really a dissolving of the other metal by the mercury, much as water dissolves sugar. The amalgam with which the dentist fills a cavity in a tooth consists commonly of tin, silver, and mercury. They are mixed just before using to a stiff paste, which hardens to an insoluble substance in a few hours. A tin amalgam was formerly used to coat the backs of mirrors. In gold and silver mining, the finely-ground ore is extracted with mercury; the latter dissolves the particles of gold and silver and leaves the quartz and other residues behind.

Compounds of Mercury.—Mercury forms two chlorides: *mercurous chloride,* HgCl, called calomel, is considerably used in medicine as a purgative and as a stimulant for the secretory organs; *mercuric chloride,* $HgCl_2$, called corro-

sive sublimate, is a powerful poison, and as such finds good use as a disinfectant for surgical instruments and the hands of surgeons, for seed-potatoes to kill diseases, and for museum specimens. *Mercuric fulminate* is the explosive in percussion caps, used to discharge other explosives.

Silver.—Silver is a white metal, takes a very high polish, corrodes rather quickly, especially in atmosphere contaminated with sulfur fumes from the burning of soft coal, and is softer than copper, but harder than gold. For this reason 100 parts of copper are added to 900 of silver for use in U. S. coins, jewelry, ornaments, table cutlery, etc; this alloy is *coin silver*. A current of electricity decomposes the salts of silver and deposits a coating of silver on one of the electrodes; knives and other objects are *silver plated* by immersing them in this solution and causing the silver to be deposited on them. Depending upon the thickness of the coat, it is called single, double, or triple plated. The backs of mirrors have some silver mixed with mercury. For coins and for ornaments, silver has been used since earliest times, as it occurs in the uncombined state in many places, and hence requires no smelting. It was the silver and the gold of the ancient Mexicans that attracted the early Spanish voyagers to our shores.

Photography.—One of the most important uses of compounds of silver is in photography. The photographic film or plate is covered with a thin film of gelatin in which is suspended a very fine powder of *silver bromide,* AgBr, or *silver chloride,* AgCl. When the plate is put into a "developer," the latter decomposes the silver bromide rapidly wherever the light struck it, but very slowly where it did not. In the decomposition, black particles of silver are deposited. Thus, on developing an exposed plate, the parts exposed to the most light will develop black, and

those exposed to the least light will develop gray or remain white. This constitutes a "negative," since the lighter parts of the object photographed appear dark, and the darker parts appear light. On laying a piece of photographic paper on the negative and allowing light to pass through the latter to the paper, a "positive" can be developed, as the darker parts of the negative will appear as the lighter parts in the paper. The science of photography has developed at an enormously rapid pace during the last thirty years. When films could be made which were so sensitive as to require only one-hundredth of a second exposure to light, pictures of a moving object could be taken at the rate of sixteen a second; and when these pictures are shown on a screen at the same rate, we have our moving picture.

Since silver compounds are poisonous, a few of them are used as disinfectants. 'Argyrol is one of these.

Gold.—Although interesting enough in a financial way, gold is chemically a rather uninteresting element, since it is very inert, forming but few compounds, and these being of no special importance. Most of the gold in nature is uncombined; hence it has been used from the earliest times. Because it does react with so few substances, however, and does not corrode in air, it finds many valuable uses as a protective coating over other metals, especially in architectural work, and in lining the vessels in which corrosive chemicals are used. Gold is extremely malleable; one grain of it can be beaten into a sheet that will cover half a square yard. This is gold leaf or gold foil. Since gold is very soft and wears away readily, it is alloyed with copper for most purposes. In jewelry, pure gold is called 24-carat; 16-carat gold means 16 parts of gold alloyed with 8 parts of copper. Due to new processes of obtaining gold from its ores, the production of this metal has increased very rapidly during the last few dec-

ades; as a result it has become much more plentiful. However, since gold is our standard of value, instead of saying that gold has become cheaper, we say that other commodities have become more expensive.

Platinum.—This is a silvery-white metal, very resistant to the action of most chemicals. For this reason it is much used as the lining of machinery employed in making nitric,

TABLE VIII

Summary of the Properties of the Heavy Metals

Metal	Where found and in what principal form	Uses of the metal and its compounds
Iron......	As Fe_2O_3, hematite As Fe_3O_4, magnetite	Fe—most important of all metals industrially; used as wrought-iron, cast-iron, and steel. $FeSO_4$—killing weeds; in dyeing; in making ink. Fe_2O_3—paint pigment. Iron is an essential element in both plant and animal life; occurs in red blood corpuscles.
Lead......	As PbS, galena	Pb—used for water pipes and in batteries; in alloys, as babbitt and solder. $PbSO_4$—paint pigment. $Pb(OH)_2.2PbCO_3$—white lead, paint pigment.
Copper....	Found in large quantities as Cu As CuS	Cu—used for electric wire, roofs, in many alloys, as coins, brass, bronze, gun-metal $CuSO_4$—insect poison.
Tin........	As SnO_2	Sn—used in tin plating; in alloys, as bronze, solder, pewter.
Zinc......	As $ZnCO_3$ and ZnS	Zn—used in galvanizing iron; in batteries; in alloys, as brass. ZnO—paint pigment; medicine.
Mercury...	As HgS	Hg—used in thermometers and barometers. HgCl—calomel; medicine. $HgCl_2$—corrosive sublimate; poison, disinfectant.
Silver.....	As Ag and AgS	Ag—used for coins, jewelry, plating. AgCl and AgBr—photographic plates and papers.
Gold......	As Au	Au—used for coins and jewelry. No important compounds.
Platinum..	As Pt	Pt—used in jewelry and electrical instruments. No important compounds.

sulfuric, and other strong acids. It is used in jewelry for setting gems, as it is hard, tough, and non-corrosive; and as the contact points of electrical instruments, because it is not oxidized by the sparks. It is our heaviest metal, being 21.47 times as heavy as water. Like gold, its compounds are rare and unimportant. It is our most costly metal.

QUESTIONS

1. Why was iron not much used by the ancients?
2. Besides abundant iron ore, what other natural product must be cheap and abundant before iron can become cheap?
3. Describe the chemistry of the blast furnace, including both the iron and the slag.
4. What are the chemical differences between wrought-iron, cast-iron and steel?
5. Name some important alloys of iron.
6. Name five compounds of iron and state their uses.
7. Why is it safer to use lead pipes with hard water than with soft?
8. What is white lead, and how is it made?
9. State three ways in which the poisonous properties of lead and its compounds may become important.
10. Why could copper, tin, gold, and silver be utilized by man from the earliest times?
11. What use is made of the poisonous nature of copper compounds?
12. Why are zinc, tin, and silver used to plate iron?
13. What is bronze? Solder? Dentist amalgam? Sterling silver? Brass? Galvanized iron?
14. What is calomel? Corrosive sublimate? Blue stone? Chlorophyll?
15. Describe the chemistry involved in making a photograph.

LABORATORY EXPERIMENTS

39. To Show the Chemistry of Iron Rust.—Force a wad of steel wool or a muslin bag full of iron filings into a large test tube or graduated cylinder so that it remains suspended half way. Pour water into the tube to the depth of an inch, wetting the iron. Invert the tube in a dish containing about an inch of water, and support it by means of a clamp. Allow to stand for a few days, noting any change in the level of the water inside the tube, and any temperature effect on the tube near the mass of iron. If the water rises an inch or more in the tube, remove the latter and test the air in it with a glowing splinter. What kind of a chemical change is the rusting of iron? Why does painting prevent iron from rusting?

40. To Reduce Copper From an Ore.—Grind together equal volumes of copper oxide and charcoal. Place the mixture in a test tube. Clamp the tube in an almost horizontal position, so as to spread out the mixture in a thin layer. Carefully heat, at first gently, then as hot as the tube will

stand, for several minutes. A wing-top burner is best. Cool, and note the appearance of the mixture. What has been formed?

$$CuO + C = Cu + CO$$
copper oxide   carbon   copper   carbon
monoxide
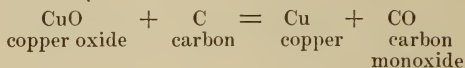
This same method is used for the reduction of many ores, especially iron. Coke is the usual form of carbon. What use is made of the CO? (Fig. 35).

41. To Reduce Mercury From an Ore.—Heat some mercuric acid in a test tube. Note the drops of metallic mercury that form on the sides of the tube.

$$HgO = Hg + O$$

Test for any increased amount of oxygen in the air in the tube.

42. To Show the Reduction of Silver Nitrate by Sunlight.—(a) Moisten a strip of filter paper with a solution of silver nitrate, then expose to direct sunlight. What is the dark substance formed? What great use is made of this property of silver salts?

(b) In what different ways can metals be separated from their compounds as illustrated by the last three experiments?

SOME COMMON NON-METALS

Significance of Metals and Non-metals.—We have seen in the preceding chapter that among the metallic chemical elements are some of our most important substances in everyday use. Many of the metals that we discussed, such as sodium, calcium and potassium, are not used as metals, but their compounds are of the utmost necessity not only to processes outside of our body, but to our very life processes themselves. Other metals, as gold and platinum, find practically all their usefulness as elements and none as compounds. Copper, iron, and lead are important both as elements and in compounds.

It may have been noticed that most of the compounds of the metals were salts; that is, the metal had replaced the hydrogen of an acid to form a salt. If the reader looks back through the last two chapters at the various salts discussed, it will be noticed that the majority of the salts are chlorides, sulfates, nitrates, phosphates and carbonates. That is, there are about half a dozen acids that form the most important salts with all metals. The significance of that is this: the non-metallic elements always form the acid radicles; the non-metallic elements involved in the above salts are, of course, Cl, S, N, P, and C; therefore, since their salts appear to constitute the important compounds of the metals, these non-metals should now occupy our attention.

Hydrogen and Oxygen as Non-metals.—Besides the five non-metallic elements just mentioned, we have already discussed *hydrogen* and *oxygen*. It might be well to point out here that these two non-metals are an exception to

the statement that the non-metals form acid radicles. Oxygen may be a part of the radicle, as in the sulfate radicle, $-SO_4$; but in the chlorides and iodides it is not. Oxygen is never the radicle itself. Hydrogen is the characteristic element of all acids; it is the presence or absence of a hydrogen replaceable by a metal that determines whether or not a compound is an acid. The fact that hydrogen and oxygen do not form acid radicles of their own is not the important thing, however; the fact of greatest importance is, *that hydrogen is present in all acids, that the OH group is present in all alkalies, and that whenever a salt is formed the hydrogen and the OH always form water.* Thus we can consider the non-metals, hydrogen and oxygen, as being in a class by themselves. They are certainly not metals, for they are gases; but, nevertheless, they do not form acids as other non-metals.

We can see, then, that the study of a non-metal involves first, the study of the element alone; second, the acid or acids which it forms; and third, the salts of these acids. In some cases there are important oxides as well; but the oxides, as we shall see, are one phase of the acids.

The most important and commonly occurring non-metallic elements are hydrogen, oxygen, nitrogen, carbon, sulfur, phosphorus, chlorine, and silicon. To this might be added the less important elements, iodine and arsenic.

Nitrogen.—Nitrogen, as we have seen, is the chief gas of the air. It is inactive there, not combining with other elements under ordinary circumstances, and not being utilized by plants or animals, with the exception of the legumes. Under the influence of an electric discharge, such as lightning, nitrogen and oxygen combine to form an oxide, which dissolves in the descending rain water to form nitric acid, HNO_3:

$$2\,N \quad + \quad 5\,O \quad = \quad N_2O_5$$
$$N_2O_5 \quad + \quad H_2O \quad = \quad 2\,HNO_3$$
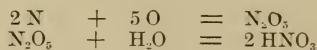
This is one of the ways in which nitric acid is formed in nature; the most important way, in the soil, we shall consider in a later chapter.
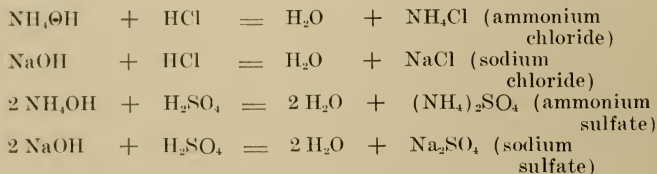
Nitric Acid.—This is the most important acid which nitrogen forms; it forms salts with practically all metals, some of which are of great importance. In the soil, nitric acid exists as the nitrates of sodium, potassium, calcium, and magnesium. Two of our commonest nitrogenous fertilizers are *sodium nitrate* (Chile saltpeter) and *calcium nitrate*. The latter is made artificially, as we have seen, by an electrical process. Since all nitrates are very soluble in water, they are easily leached out of soils; this is the reason why there is seldom any considerable accumulation of them in soils, and why nitrate fertilizers are not added once in several years, but several times in one year, in small amounts, so that the plants can use them up before they are leached out.

Potassium nitrate, KNO_3, is commonly used in gunpowder and in fireworks; it finds some use also in the curing of meat, as in dried beef and corn beef.

Ammonia.—One of the most remarkable facts about the chemistry of nitrogen is that with oxygen it forms a strong acid, and that with hydrogen it forms an alkali. The latter is ammonia, NH_3; it can be made artificially, and it is always formed when manure ferments; but practically all of our commercial ammonia is a by-product of the coal-gas works. Ammonia is an invisible gas, but it has a very penetrating, biting odor. It is soluble in water to such an extent that one quart of water will dissolve 700 quarts of ammonia gas. This water solution is our ordinary household ammonia water, or "aqua ammonia," or "spirits of hartshorn." A chemical reaction really takes place when the gas dissolves in water:

$$NH_3 \ + \ H_2O \ = \ NH_4OH$$

9

This compound, ammonium hydrate, shows the typically alkaline characters of hydrates: it turns litmus paper blue, and neutralizes acids, forming ammonium salts and water in a way exactly analogous to that of the hydrates of metals:

$$NH_4OH + HCl = H_2O + NH_4Cl \text{ (ammonium chloride)}$$
$$NaOH + HCl = H_2O + NaCl \text{ (sodium chloride)}$$
$$2 NH_4OH + H_2SO_4 = 2 H_2O + (NH_4)_2SO_4 \text{ (ammonium sulfate)}$$
$$2 NaOH + H_2SO_4 = 2 H_2O + Na_2SO_4 \text{ (sodium sulfate)}$$

Some of these ammonium salts are of great practical importance. *Ammonium sulfate* is one of our commonest nitrogenous fertilizers; *ammonium chloride,* sal-ammoniac, is much used in electric batteries, in medicine, and in dye-works. The old-fashioned smelling-salts were *ammonium carbonate.* This compound gradually breaks down into ammonia:

$$(NH_4)_2CO_3 = NH_3 + NH_4HCO_3 \text{ (ammonium bicarbonate)}$$

The inhaled ammonia gas exerts a mildly stimulative effect on the nerves.

Refrigeration with Ammonia.—Ammonia gas can be readily liquefied by pressure in steel cylinders. When the compressed gas is allowed to escape into a chamber and expand, it absorbs a great amount of heat, reducing the temperature of the chamber very considerably. This phenomenon is utilized in refrigerating plants and artificial ice factories (Fig. 38). The cooling chamber is a series of tubes surrounded by strong brine. The brine becomes very cold, far below the freezing-point of water, without itself freezing. This cold brine is circulated around small tanks of water, and the water gradually freezes to ice. Or the brine is piped around the walls

of cold-storage rooms to cool the air.   The expanded
ammonia gas is again liquefied, and thus is used over
and over again.

Explosives.—A very noteworthy characteristic of many
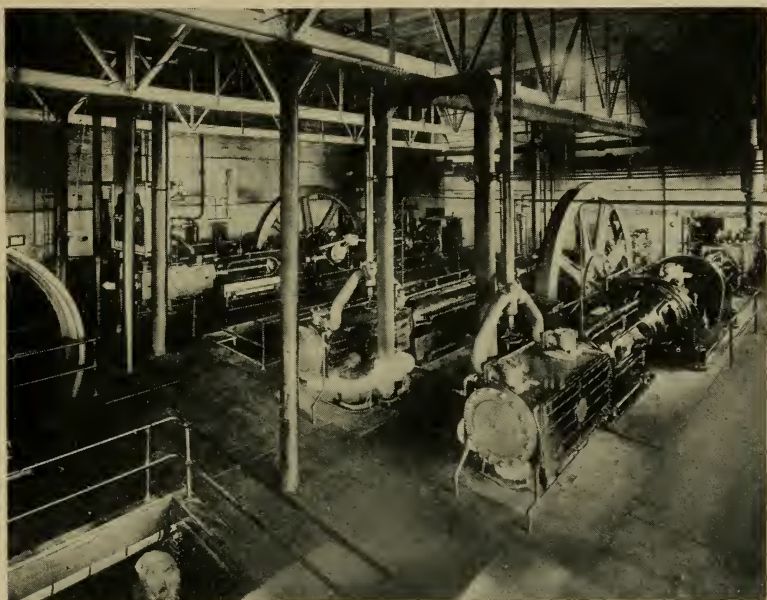compounds of nitrogen is that they decompose when

FIG. 38.—Scene in a refrigerator room of a large packing house.   Bacteria cannot thrive at
low temperatures; hence cold storage is a very effective means of preservation of meat, eggs,
dairy products and fruit.   The cold is produced by compressing ammonia gas with the ma-
chines shown here, then allowing the gas to expand in chambers surrounded by brine.   The
expansion cools the gas, which cools the brine, which is then piped to the cold storage rooms.

jarred or ignited, and liberate large volumes of gaseous
products.   These unstable compounds are our high ex-
plosives.   Some of them are nitroglycerin, nitrocellulose
(gun-cotton or smokeless powder), and trinitrotoluol
(the famous T. N. T. of the World War).   Celluloid and
collodion are near chemical relatives to nitrocellulose.

Nitrogen in Living Cells.—We shall have occasion in later chapters to discuss the relation of nitrogen compounds to living animals and plants. Nitrogen compounds constitute the living material of all plant and animal cells. This is such an all-important subject that it demands special attention.

Carbonates.—The next non-metallic element, *carbon*, we have already discussed at some length. We saw that as an element it is our most important solid fuel; that combined with hydrogen it forms our principal liquid and gaseous fuels; that combined with hydrogen and oxygen it forms thousands of different compounds; that it has two important oxides, the monoxide and the dioxide; that the latter occurs in small amounts in the air, and furnishes all green plants with their source of carbon; that enormous amounts of carbon dioxide are held in the earth in the form of limestone.

Carbon dioxide, when dissolved in water, forms a solution that has a slightly sour taste, and will gradually redden litmus paper. It must therefore form an acid:

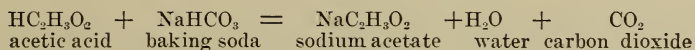$$CO_2 \;+\; H_2O \;=\; H_2CO_3 \text{ (carbonic acid)}$$

It is not a very well-defined acid, but it does form a series of salts with many metals. These salts are the *carbonates*. We have already mentioned the carbonates of sodium, potassium, calcium, magnesium, and of several heavy metals. We have seen that *sodium carbonate,* Na_2CO_3, is our sal-soda, or washing soda; that *sodium bicarbonate,* $NaHCO_3$, is our baking soda or saleratus; that calcium carbonate occurs as limestone, marble, and chalk; that potassium carbonate is the form of potash in wood ashes. We have also emphasized a very important chemical reaction which all carbonates show, and that is that *they give off carbon dioxide when treated with an acid*. In order to

understand the full significance of this reaction, let us take, for example, sodium carbonate and hydrochloric acid:

$$Na_2CO_3 \; + \; 2\,HCl \; = \; 2\,NaCl \; + \; H_2O \; + \; CO_2$$
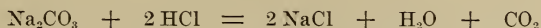
We start with a strong acid; the products formed are common salt, water, and carbon dioxide, which escapes as a gas. In other words, we destroyed an acid, and have left only a neutral salt. That is, *the carbonate in reality neutralized the acid*. Thus, if we have some vinegar, and carefully add baking soda to it, the soda, being a carbonate, will react with the acetic acid of the vinegar, and carbon dioxide gas will bubble up out of the vinegar and escape into the air. If we add just enough soda, all of the acid and all of the soda will be destroyed, and there will be left behind a neutral vinegar, containing a little sodium acetate, a salt, in solution:

$$HC_2H_3O_2 \; + \; NaHCO_3 \; = \; NaC_2H_3O_2 \; + H_2O \; + \; CO_2$$
acetic acid   baking soda   sodium acetate   water   carbon dioxide

Carbonates thus apparently constitute a special form of alkali, since they neutralize acids to form salts and water. The keynote to the whole reaction is simply the fact that carbon dioxide is always formed; this is a weak acid, and thus takes the place of the acid neutralized; and as it is a gas, it escapes and removes itself from the reacting substances entirely. There are many places where carbonates are used to neutralize acidity: in cooking with sour milk or molasses, baking soda is used; when a soil becomes acid, limestone is added to it; in the preservation of certain fruit juices and in the manufacture of some sugar products from acid juices, limestone is often used.

Sulfur.—This is probably one of the oldest-known chemical elements. It is mentioned often in the Bible,

usually being called "brimstone." In some way or other
the choking fumes of burning sulfur suggested that the
fires of Hades must be concerned with this substance, and
hence brimstone has always been coupled with Hades in
the superstitious mind. Sulfur is found as the element
in many places in the earth, especially in volcanic regions.
In the Japan Archipelago there are whole islands of fairly
pure sulfur; in Sicily, Spain, California, Mexico, and
Yellowstone Park, there are large deposits of sulfur that
no doubt had their origin in volcanoes. In Texas and
Louisiana there are immense beds of sulfur, several hun-
dred feet below the ground, that had an entirely different
origin. There are certain bacteria and algæ which decom-
pose sulfates as gypsum, $CaSO_4$, and deposit minute
granules of sulfur in their cells. After the death and the
decay of the cells, the sulfur granules are left, and grad-
ually accumulate in beds, similar to the formation of lime-
stone beds. These sulfur beds thus furnish another
interesting connecting link between the organic or living
world and the inorganic or mineral world.

Commercial Sulfur.—The mining of sulfur when it
occurs above ground offers no more difficulties than the
digging of sand. In this country, however, most of the
sulfur comes from the underground beds of Texas and
Louisiana and a very neat bit of chemical engineering is
involved in obtaining this sulfur (Fig. 39). Instead of
mining it like coal, a far cheaper method is used, consist-
ing of melting the sulfur by superheated steam and then
pumping the liquid sulfur to the surface. As sulfur melts
at about 245° F., it is easily melted by the steam. The
molten sulfur is poured into moulds to solidify, or into
huge tanks from which the solid sulfur is removed by
blasting. In this form it is called commercially "roll-
sulfur" or "brimstone." At a high temperature sulfur
boils; if the vapors are cooled and condensed on the sides

of brick vaults, they form a finely-powdered, crystalline sulfur called "flowers of sulfur."

Uses of Sulfur.—Something like half a million tons of sulfur are used in the United States every year. Both the element and its compounds are important. Flowers of
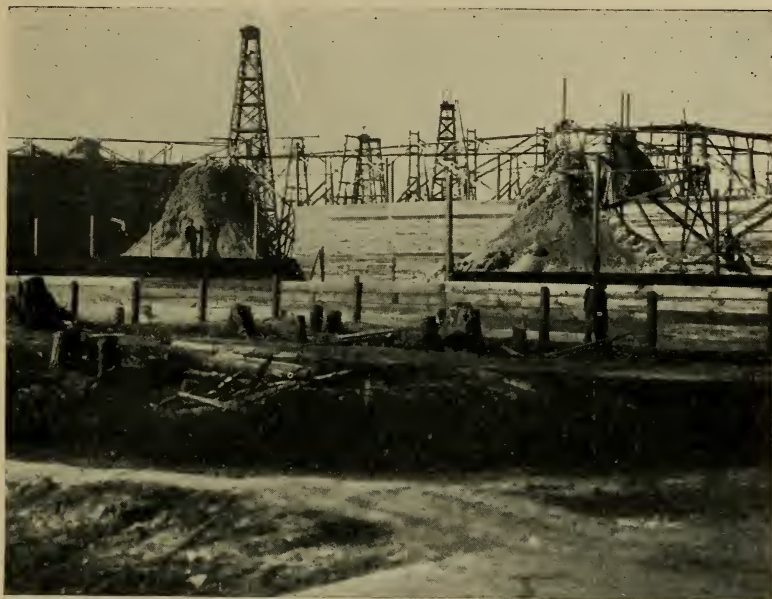

FIG. 39.—Texas sulfur wells. A hole is drilled down into the bed of sulfur, and a double pipe, one within the other, is let down the hole. Live steam is passed into the sulfur through the inner tube; the sulfur is melted and forced to the surface of the ground through the outer tube. It is pumped into huge vats, shown in the picture, where it cools, and is then ready for market. (Union Sulfur Co.)

sulfur form one of the constituents of black gunpowder; they are commonly found in veterinary patent medicines; they are sometimes used in poultices in human medicine; and they are used very considerably in the vulcanizing of rubber. When it is burned, *sulfur dioxide* is formed.

$$S + 2O = SO_2$$
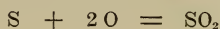
This is an acrid, biting, choking gas. It has two very valuable properties, its *bleaching action* and its *germicidal effect*. Immense quantities are used in bleaching straw, paper, linen, dried fruits, molasses, and other products. As a disinfectant it is often burned in chicken houses, barns, and residences to kill disease germs and insect pests. Sulfur dioxide gas is now compressed into steel cylinders for use in disinfecting and bleaching. Another use of sulfur is in the combination of slaked lime and sulfur, called "lime-sulfur," used as a dip for sheep and hogs, and for killing scale insects on trees.

The King of Chemicals.—By far the greatest use of sulfur is in the manufacture of *sulfuric acid,* H_2SO_4. This acid is not in itself of such great importance; but there is hardly a chemical industry of any sort which does not utilize sulfuric acid at least indirectly to prepare its products. We have already seen that hydrochloric and nitric acids are prepared from their salts by means of sulfuric acid; it is the mechanism by which nitric acid is made to combine with glycerin, cotton, and toluol in producing our high explosives; it is used in coal-gas factories for washing the ammonia from the gas; our acid phosphate fertilizers are formed by the action of sulfuric acid on bones or rock phosphate; the manufacture of most of our drugs, dyes, cosmetics, pure metals, and hosts of other commodities, involves the use of sulfuric acid somewhere in the processes. Because of this extraordinary usefulness, sulfuric acid is often termed "the king of chemicals."

Sulfides and Sulfates.—Sulfur occurs very commonly in nature combined with metals as *sulfides*. Fools' gold is a sulfide of iron; the principal zinc and tin ores are sulfides. The salts of sulfuric acid are also rather abundant. *Potassium* and *magnesium sulfates* are found in

the Stassfurt potash mines; *magnesium sulfate,* we have seen, is a common constituent of so-called alkali waters; there is some *sodium sulfate* in ocean waters; the alabaster of Biblical literature is a certain form of *calcium sulfate.* Sulfur compounds are found in all soils, and in all animals and plants. In fact, *it is one of the necessary elements for life processes, both plant and animal.*

When eggs decay, the peculiarly offensive odor is due to *hydrogen sulfide,* H_2S. The characteristic odor of mustard, onion, garlic, and cabbage is due to certain organic compounds of sulfur. One of the most vicious gases used by the Germans in the World War is called ''mustard gas''; it is a very poisonous compound of sulfur, but it is not the compound occurring naturally in mustard.

Phosphorus.—This non-metallic element does not occur in nature as an element. The reason for this is apparent when a piece of phosphorus is exposed to the air; it immediately begins to fume, showing that it is slowly oxidizing. We have all rubbed match-heads in the dark, and watched them fume and glow; this is due to the fact that there is phosphorus in the head, and that the rubbing of it exposes fresh phosphorus to the air. Certain bacteria emit a dim light which is similar to the glow of phosphorus; hence the name of ''phosphorescence'' applied to this light, although phosphorus is not involved in the bacterial process, any more than it is in the light of fireflies and of glow-worms.

Practically all the phosphorus in nature exists as *calcium phosphate,* either in rock phosphate or in the mineral apatite, which is found in granite rocks and in all soils. It is not abundant in soils, however, and as growing plants require considerable amounts of it, phosphorus fertilizers are very much in demand.

Yellow and Red Phosphorus.—There is one peculiarity about phosphorus that makes it unique among all other

elements, and that is that elemental phosphorus exists in two forms. The ordinary phosphorus is a yellow, wax-like, semi-transparent solid. When it is heated to a certain temperature away from air it is converted into a brick-red solid. This red substance is still uncombined phosphorus, but it now has entirely different properties from what it had before. The *yellow phosphorus* catches on fire at about 100° F.; the red at about 400° F. The red does not glow in the dark. Yellow phosphorus is exceedingly poisonous, while the red is not. The yellow dissolves in ether, the red does not.

Uses of Phosphorus.—Yellow phosphorus is used in some rat poisons. These are dangerous to use around buildings, however, as the phosphorus very often ignites in warm weather, and starts a' conflagration.

Matches.—The most important use of phosphorus is in *matches*. Advantage is taken of the fact that phosphorus catches on fire at such a low temperature that the heat of friction will ignite it. The head of the match consists of (1) phosphorus, which does the initial igniting; (2) potassium chlorate, which readily liberates large amounts of oxygen to continue the burning started by the phosphorus; (3) fine sand, clay, or powdered glass to increase the friction; and (4) glue, to bind these materials into a solid head. Scratching on any surface will ignite these matches. Double-dip matches have phosphorus only on the point, thus lessening the danger of their being ignited by accident. A still further improvement is the safety match, which bears no phosphorus on the head; it is ignited by scratching on the side of the box, where the phosphorus and powdered glass are glued. Formerly yellow phosphorus was the form used, as it ignites more readily. The poisoning of workers in match factories, as well as of children who accidentally chewed the heads of the matches, has brought about legislation in most coun-

tries which demands the use of the non-poisonous red variety.

Phosphates.—Practically the only compounds of phosphorus which are of importance are the salts of phosphoric acid, which has the formula H_3PO_4. We saw on
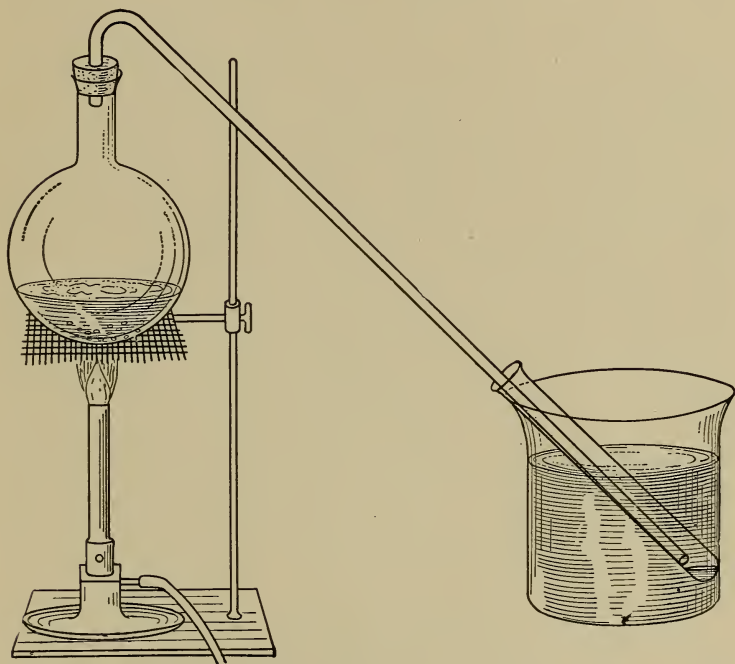

FIG. 40.—Apparatus for the preparation of nitric acid.

page 84 how this acid is formed when the fumes of burned phosphorus dissolve in water.

The calcium salt, $Ca_3(PO_4)_2$, is found in large quantities as rock phosphate, or "floats," in our Southern States, and is our largest source of phosphate fertilizer. A great deal of it is converted into "acid phosphate," $CaH_4(PO_4)_2$, which is used as a soluble phosphate ferti-

lizer, and in the purification of juices in sugar manufacture. The bones of animal bodies consist almost entirely of calcium phosphate, which fact suggests the possible origin of the rock phosphate beds.

Chlorine.—This gaseous element does not occur in nature as the element but is very abundant in the earth in the form of chlorides. The abundance of sodium chloride in the earth and in the sea has been mentioned before; calcium, potassium, and magnesium chlorides are also important in the Stassfurt mines. The element chlorine is a greenish-yellow gas, exceedingly irritating to the mucous membrane of the nose and throat, causing severe "chlorine colds." In even small amounts it is very poisonous. One of the first materials used by the Germans when they commenced their gas warfare was chlorine. The sufferings of the attacked soldiers were terrible. The gas masks devised as a protection against the chlorine contained various chemicals which converted the chlorine into harmless salts. The principal uses of chlorine are in bleaching cotton, straw, linen, and other materials, and as a disinfectant in water. For these uses, either the element itself stored in steel cylinders, or the compound called "hypochlorite" or "bleaching powder," is used (Fig. 7).

Compounds of chlorine are found in all plant bodies, although it is probably not a necessary element there. In the animal body, however, it is necessary, as the acid of the stomach juice is *hydrochloric acid*, HCl. This is the "muriatic acid" of the tinner, used in soldering.

Silicon.—Our interest in the compounds of *silicon* is mainly concerned in their very abundant occurrence in nature, three-fourths of the weight of the earth being *silicon dioxide*, SiO_2. This substance constitutes all of sand and sandstones, and the greater part of most other

rocks except limestone. The *silicates* have rather complex formulas. We shall simply mention that some of the more useful silicates are *talc, pumice,* and *asbestos;* that the *aluminum silicates* constitute all our clays; that Portland cement is a mixture of lime and aluminum silicate; and that "water-glass" for preserving eggs is *sodium silicate.*

Arsenic and Iodine.—There are two other non-metallic elements that have important special uses. These are *arsenic* and *iodine.* The first is a very poisonous element in all its compounds; hence it is utilized in most of our insect and vermin poisons, such as Paris green, white arsenic, lead arsenate, and calcium arsenate. Iodine as an element is a good disinfectant; it kills germs readily, but is not irritating or caustic to raw flesh. In wounds, on bruises, and on tonsils it is very effective. In the form of potassium iodide it is often taken as a medicine, especially for goitre. The compound called *iodoform* is a common disinfectant.

Oxides and Non-metals.—In our discussion of the kinds of compounds that are formed by the nonmetals on page 128 we mentioned that although acids and salts were

Fig. 41.—Apparatus for the preparation of ammonia.

the principal ones, some oxides were also important. Let us now see exactly the relation which oxides bear to acids. For example, let us take a few of the non-metals, and, in
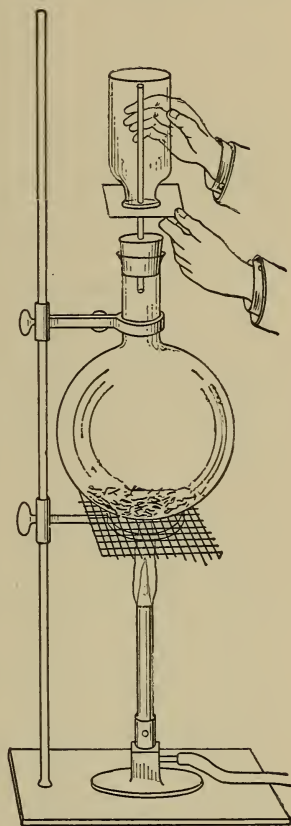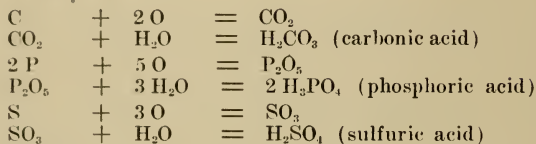
equations, show how they form oxides, and then how these oxides unite with water.

$$C \quad + \quad 2\,O \quad = \quad CO_2$$
$$CO_2 \quad + \quad H_2O \quad = \quad H_2CO_3 \text{ (carbonic acid)}$$
$$2\,P \quad + \quad 5\,O \quad = \quad P_2O_5$$
$$P_2O_5 \quad + \quad 3\,H_2O \quad = \quad 2\,H_3PO_4 \text{ (phosphoric acid)}$$
$$S \quad + \quad 3\,O \quad = \quad SO_3$$
$$SO_3 \quad + \quad H_2O \quad = \quad H_2SO_4 \text{ (sulfuric acid)}$$

It will be seen that the oxides of the non-metals combine with water to form the corresponding acids. This fact has its important applications. For example, if sulfur dioxide gas, SO_2, is being used as a bleaching agent, the material to be bleached must be moist, since the bleaching effect is caused by sulfurous acid, formed thus:

$$SO_2 \ + \ H_2O = H_2SO_3 \text{ (sulfurous acid)}$$

Again, when our soft drinks are charged with carbon dioxide, the drinks have a sour taste because the carbon dioxide forms carbonic acid with the water.

QUESTIONS

1. What is the most important characteristic distinguishing metals from non-metals?
2. Name the non-metallic elements we have studied, and state which of them form important acids.
3. Explain the relation of hydrogen and oxygen to acids and alkalies.
4. Name the six most important compounds of nitrogen, and give the use of each.
5. Explain the action of a carbonate in neutralizing an acid.
6. What is carbonic acid? Name three of its important salts, and give the uses of each.
7. Explain in what ways deposits of carbon and of sulfur in the earth form the connecting link between the organic and the inorganic world.
8. Explain the important uses of elemental sulfur.
9. What are some of the most common uses of sulfur in the industries? On the farm?
10. Explain the chemistry of a safety match.
11. What is the most abundant compound of phosphorus in our bodies?
12. What is chlorine gas good for?
13. Name some chlorides and give their uses.
14. What is the most abundant element in the earth? The most abundant compound?
15. What is Portland cement? Water-glass? Paris green?
16. Explain the relation between the oxides of the non-metals and the acids of these elements.

TABLE IX—*Summary of the Properties of the Non-metals*

| Element	Where found and in what form	Uses of the element and its compounds
Hydrogen.	Principally in H_2O	H is used for welding; for filling balloons. H is a constituent of all acids. Enters into large number of important compounds.
Oxygen....	As O in air As oxide of a great many elements in nature	O supports combustion. Compounds too numerous to mention.
Nitrogen..	As N in air Compounds of N very scarce in nature	N dilutes the oxygen of the air. NH_3—fertilizer; cleaning agent. HNO_3—the acid of many important nitrates, as $NaNO_3$, KNO_3, $Ca(NO_3)_2$ Most explosives are compounds of N. All plants and animal cells contain compounds of N.
Carbon	(See end of Chapter VI.)	
Sulfur.....	In great quantities as the element As FeS_2, pyrites, or "fool's gold" As K_2SO_4 and $MgSO_4$ in the Stassfurt mines	S—used in medicine; in vulcanizing rubber; in gunpowder. SO_2—bleaching straw, sugar cane juice, and fabrics; disinfecting. H_2SO_4 is the most important single chemical manufactured; in the making of explosives, fertilizers, dyes, drugs, and many other commodities. Compounds of S are essential to both plant and animal life.
Phosphorus	As $Ca_3(PO_4)_2$, rock phosphate	P—used in match-heads; in rat poisons. Bones consist of $Ca_3(PO_4)_2$. Na_2HPO_4 used in medicine. Compounds of P are essential to both plant and animal life.
Chlorine...	Principally as NaCl, KCl, and $MgCl_2$ In sea-water and in all soils	Cl—used in bleaching; in disinfecting water. $CaOCl_2$—bleaching powder. HCl—in the stomach; is muriatic acid. NaCl—most abundant compound of Cl. Compounds of Cl are essential to animal but not to plant life.
Silicon	As SiO_2 to the extent of three-fourths of the earth's crust	Si—no important uses. SiO_2—sand; found in most rocks and minerals. Talc, mica, asbestos, pumice are complex silicates. Portland cement is aluminum silicate. Water glass is sodium silicate.
Arsenic....	In connection with copper and zinc ores	As_2O_3-white arsenic; insect poisons and sprays; in the manufacture of Paris green and lead arsenate.
Iodine.....	In sea-weeds	I—used as disinfectant. CHI_3—iodoform; disinfectant.

LABORATORY EXPERIMENTS

43. To Prepare Nitric Acid.—(To be done by instructor.) Set up the apparatus shown in figure 40, using a rubber stopper. A glass retort is preferable, as the acid quickly attacks the rubber. The receiving test tube stands in cold water. Place 25 g. sodium nitrate in the flask, then add 14 c.c. of concentrated sulfuric acid and fit in the stopper. Gently mix the contents, and then carefully heat with a low flame over the nitric acid. Save for experiment 44.

$$2\,NaNO_3 \;+\; H_2SO_4 \;=\; 2\,HNO_3 \;+\; Na_2SO_4$$

44. To Study the Properties of Nitric Acid.—Use the acid prepared in experiment 43, or that in the reagent bottle.

(a) Test it with litmus paper.

(b) Dilute a little with water, and add a carbonate.

(c) Add a few drops of the concentrated acid to wool, finger-nail clippings, egg white, and other forms of animal tissues. This yellow color is characteristic of the nitrogenous substances called *proteins*.

(d) Into one test tube put a few c.c. of nitric acid diluted with an equal volume of water; into another put some diluted sulfuric acid; into another, hydrochloric acid. Then add a small piece of copper wire to each tube, and warm gently. Test any evolved gas for hydrogen. Compare the results with those in experiment 12. Do all metals displace the hydrogen from acids? What gas is evolved from nitric acid by the copper (see equation below)? To what is the blue color due?

$$3\,Cu \;+\; 8\,HNO_3 \;=\; 3\,Cu(NO_3)_2 \;+\; 4\,H_2O \;+\; 2\,NO$$
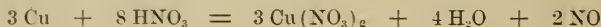
45. To Prepare Ammonia.—Set up the apparatus shown in figure 41. Place 10 g. ammonium chloride, 10 g. powdered quicklime, and 20 c.c. water in the flask. Heat gently, filling two bottles and a test tube with the gas, keeping the bottles bottom up on the table. Equation.

(a) Set the inverted test tube full of ammonia into a beaker of water. Why does the water rise in the tube? What is ammonia water?

(b) Hold moist strips of litmus paper in one of the bottles. What does this show?

(c) Moisten a glass rod with concentrated hydrochloric acid, then hold it over a bottle of ammonia. What are the white fumes?

$$NH_3 \;+\; HCl \;=\; NH_4Cl$$

46. To Study the Properties of Sulfur Dioxide.—(a) What was learned concerning the chemistry of sulfur dioxide in experiment 9? Write equations showing the formation of sulfur dioxide.

(b) Suspend a piece of wet colored calico, some moist straw, and a pink flower in a large inverted bottle. Set the bottle over a 2 g. heap of burning sulfur on a sheet of asbestos or on a shallow iron plate. Examine the next day. For what is sulfur dioxide used commercially? How is it handled?

47. To Study the Properties of Sulfuric Acid.—(a) Prepare a 10 per cent solution of sulfuric acid carefully pouring 3 c.c. (5 g.) of concentrated sulfuric acid into 45 c.c. of water in a beaker. By means of a fine

glass rod write on a piece of paper with the acid, and dry over a flame. To what is the brown coloration due?

(b) Pour some concentrated sulfuric acid into a little moist sugar in a beaker. The beaker should stand on a large sheet of paper. What is the black substance?

(c) Test the solubility of sulfates of Na, K, Mg, Ca, Cu and Fe, by the method used in experiment 13. To a dilute solution of one of the sulfates add a little hydrochloric acid, then a few drops of $BaCl_2$ solution.
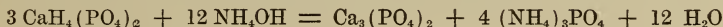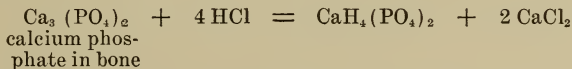
$$Na_2SO_4 \;+\; BaCl_2 \;=\; BaSO_4 \;+\; 2\,NaCl$$

The precipitate is barium sulfate, $BaSO_4$. *This is a commonly used test for sulfates.*

48. To Study Phosphorus and Its Compounds.—(a) Hold a small piece of phosphorus in tongs in the dark. What chemical reaction goes on? Why is phosphorus kept under water? How can phosphoric acid be made? Compare with experiment 8.

(b) Test the solubility of sodium phosphate, Na_3PO_4, tricalcium phosphate, $Ca_3(PO_4)_2$, and acid calcium phosphate, $CaH_4(PO_4)_2$. Which of the latter two would be chosen as the more quickly available fertilizer (p. 167)?

(c) Make a dilute solution of sodium phosphate slightly alkaline with ammonia, then add a few drops of a solution of calcium chloride. Treat some ground bones with dilute hydrochloric acid, and neutralize with ammonia. Compare the precipitate of calcium phosphate with that obtained in the first test.

$$\underset{\substack{\text{calcium phos-}\\\text{phate in bone}}}{Ca_3\,(PO_4)_2} \;+\; 4\,HCl \;=\; CaH_4(PO_4)_2 \;+\; 2\,CaCl_2$$

$$3\,CaH_4(PO_4)_2 \;+\; 12\,NH_4OH \;=\; Ca_3(PO_4)_2 \;+\; 4\,(NH_4)_3PO_4 \;+\; 12\;H_2O$$

49. To Study Chlorine and Its Compounds.—(a) Put about one-half gram of manganese dioxide in the bottom of a test tube, and add a few drops of concentrated hydrochloric acid. Warm gently. The manganese dioxide furnishes oxygen for the following reaction:

$$2\,HCl \;+\; O \;=\; 2\,Cl \;+\; H_2O$$

Test the odor of the evolved gas by wafting a little of it towards the nose by a wave of the hand. *Do not breathe the gas directly.* Hang a damp strip of colored calico in the tube by means of a loosely fitting cork, and let stand for 10 or 15 minutes.

(b) Make a suspension of bleaching powder in water, and immerse in it a piece of the colored calico used in (a). In two other test tubes put pieces of the calico into solutions of two other compounds of chlorine, such as NaCl, or $CaCl_2$. Judged by these three tests and the one in (a), what form of chlorine is evidently given off by the bleaching powder? Does the odor of the powder suggest the odor of the gas in (a)? For what is bleaching powder used? For what is chlorine gas used?

(c) Test the solubilities of the chlorides of Na, K, Ca, Mg, and Ba.

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(d) To a dilute solution of a chloride add a few drops of nitric acid, then a few c.c. of silver nitrate solution. The precipitate is silver chloride:

$$NaCl + AgNO_3 = AgCl + NaNO_3$$

This is a universally-used test for chlorides.

(e) Add a few c.c. of concentrated sulfuric acid to 2 g. of sodium chloride in a test tube. Warm gently. Compare the odor of the escaping gas to that of the vapors of a bottle of hydrochloric acid.

$$2 NaCl + H_2SO_4 = 2 HCl + Na_2SO_4$$

Hold moist blue litmus paper in the gas. Dip a glass rod into ammonia water, then hold the rod over the mouth of the test tube. Compare with the test made in experiment 45 (c).

50. To Prepare a Table of Solubilities.—(a) By means of the results of the previous experiments, fill in as many of the spaces in the table as is possible. If certain salts, as silver carbonate and magnesium nitrate, have not been worked with, leave the spaces blank. Use the following abbreviations to express the solubility:

<div align="center">

v.s. = very soluble
s.s. = slightly soluble
i. = insoluble

</div>

(b) State briefly how you would test an unknown substance to see whether it is a carbonate, chloride, sulfate or phosphate.

<div align="center">

TABLE X

</div>

	Sulfates	Phosphates	Nitrates	Carbonates	Chlorides
Na					
K					
Ca					
Mg					
Fe					
Cu					
Ag					
Ba					

THE AGRICULTURAL CHEMICAL ELEMENTS

In the preceding chapters, as we discussed each chemical element that came to our notice, we mentioned two general kinds of uses for that element or its compounds. One was its uses in the plant or in the animal body, the other was its uses outside of the body. The latter uses determine whether it occurs in our manufactured articles, our clothing, fuels, building materials, cleaning agents, medicines. In this connection we discussed some twenty-five different elements, and found that some of them are always found in plants, some are always found in animals, some are occasionally found in both of them, and some are never found in either. In this chapter we wish to get a general view of the elements which go to make up the bodies of plants and animals, as these form the basis of agriculture.

The Occurrence of These Elements.—It might be well at the outset to point out that of all the fifteen or so elements found in animals and plants, none are found there *as elements,* but only *as compounds.* Hence, when we speak of sulfur occurring in this or that part of the animal body, we do not mean uncombined sulfur, but a compound of sulfur. *Oxygen is the only free element that can be utilized by the higher plants and animals,* and this is soon combined with substances in the body in the process of combustion. Some argue that free nitrogen is utilized by the legumes; but this is not the case, since it is the bacteria on the roots of the clover, and not the clover plant itself, that absorbs the free nitrogen. And in this

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connection it should be stated that when we say animals and plants, we mean the higher animals and plants, and usually those of agricultural importance. The bacteria, fungi, and lower forms of animal life have many peculiarities in their nutrition and ways of living that would prove exceptions to the above statements.

The following classification of the chemical elements connected with agriculture will aid in getting the proper viewpoint concerning them:

1. *The elements necessary to plant life:*
    Carbon, hydrogen, oxygen, nitrogen, sulfur, phosphorus, iron, potassium, calcium, magnesium.
2. *The elements not necessary to plants, but usually found in them:*
    Chlorine, silicon, sodium, aluminum, iodine.
3. *The elements necessary to animal life:*
    The same as group 1, besides chlorine, sodium, and iodine.
4. *The elements found in the soil:*
    All the elements mentioned in the first three groups, besides several others of slight importance.

Elements in the Soil.—The soil contains all the elements that are found in either plants or animals. This is to be expected, since the animal gets its elements by eating the plants, and the plants get theirs from the soil. It must not be forgotten that plants get practically all of their carbon from the air, although a small amount is found in soils from the decay of plant material there. Soils very often contain other elements besides the fifteen enumerated above, when the rocks from which the soil originated contain them. Thus, copper, arsenic, barium, zinc, and manganese are frequently found in soils in small quantities.

Absorption by Roots.—When a plant is growing in soil,

the soil water dissolves the compounds of the soil, and the plant absorbs this solution through its roots. To a certain degree, the root can select the materials which it wants from this solution, and reject the ones it does not want. This is so only to a limited extent, however. For example, if copper sulfate is added to a soil, the plants growing there cannot always reject it altogether from the soil solution. They will up to a certain extent; and they will tolerate a certain amount of it in their tissues without being harmed; but if the amount of copper sulfate in the soil is so great that these limits are passed, the plant will be harmed by it and perhaps killed. In the same way, if a beneficial compound as potassium nitrate is found in excess in a soil, as it sometimes is in certain semi-arid regions, the plant cannot absorb just the amount it needs, but will be literally forced to overfeed, and thus be harmed. Certain plants have become accustomed to growing in salt water; these plants have not learned how to reject the unwanted amounts of salt, but their tissues have learned how to thrive in spite of the excess.

The Elements Necessary for Plants.—In the second group in the above outline, there are a number of elements which are ''not necessary to plants, but usually found in them.'' This naturally suggests two questions: Why are these elements found in plants if they are not necessary? and, how do we know that they are not necessary? The first question has already been answered: The plant roots absorb the soil water and all substances that are dissolved therein, whether the plant needs them or not. If the unnecessary elements were a part of the rocks and minerals from which the soil originated, they are naturally in the soil, and are then found also in the plants growing in that soil.

But how do we know that the plant does not need them? This has been reached by a long series of careful

experiments, in which many varieties of plants have been grown in artificial soils. The soils consisted of pure, carefully-washed sand, which is silicon dioxide, SiO_2, to which is added solutions of the salts of all the elements but one found in natural soil. That is, to one pot of sand is added all the fifteen elements above, but with sulfur omitted; in another, calcium is omitted; and so on. By this method, after many different trials by many workers, we now know what elements various plants absolutely have to have in order to grow and produce seed. From these experiments we are able to state in the above outline which elements are necessary, and which are not, in the life of the plant.

The Elements Necessary for Animals.—When it comes to deciding which elements are necessary in the life of the animal, we have a very different problem. We cannot feed the animal a series of salt mixtures with one element omitted, and expect the animal to grow. The food of the animal is plant material, or prepared products from plant materials. When a cow eats a shock of corn, she eats all the chemical elements that the corn obtained from the soil. In order to eliminate certain elements, we would have to take that shock of corn and prepare from it pure substances, as pure cornstarch and corn oil, which contain only C, H, and O, and pure corn protein, which contains only C, H, O, N, and S, and feed these to the cow, together with salts of the various other elements which we wanted to feed. It can be readily seen that this would be a next to impossible task.

Such experiments have been conducted on small animals like rats and rabbits, but most of our information along these lines is reached in another way. This is by the direct analysis of the tissues of animals. If a certain element is always formed in certain tissues, and if another element is never found in any tissue, it is pretty good evidence that the one element is necessary, and that

the other is unnecessary, to the well-being of the animal. We could not use this method in the case of plants, for they absorb everything that is in the soil, irrespective of any use they may have for it. Animals, however, have a much greater selective ability; when an element is taken into their body and it is not wanted, it is promptly eliminated, either through the intestines or the kidneys. Thus, if an animal day after day eats plant material which contains silicon, and if no silicon is found in any of its tissues, and if the silicon in the urine and fæces amounts to the same as that eaten in the food, the presumption is that that animal does not need the element silicon. From this kind of study we have been able to determine the list of elements necessary for animal life. It will be noticed that the animal requires three elements that the plant does not; it is very fortunate that nature has seen fit to have plants take up these elements from the soil, even though they be not for the plants' own use.

The Agricultural Chemical Elements.—In summary of what has been said concerning these agricultural chemical elements, we can bear in mind that there are about fifteen of these elements; that the soil furnishes the plants with all of them but carbon, which is obtained from the air; that these elements must exist in the soil as compounds soluble in water; that if any one of the ten elements essential to plant life is lacking, the soil will not support plant growth; that animals obtain all their chemical elements from the plants which they eat, and that if the plants contain elements not wanted by the animal, the latter excretes them from its body.

Origin of Living Matter.—It is interesting to note again at this place that water is the most abundant of all compounds in living organisms. Over three-fourths of the weight of all animals is water, and many plants are nine-tenths water. Also, it should be pointed out that the

fifteen chemical elements mentioned above are found in sea-water. There must obviously be some connection between these facts. Mathews, in his "Physiological Chemistry," says: "It is certainly not without significance that living matter is so watery and contains the salts of the sea. It would appear probable from this that living matter originated either in the sea itself or, perhaps, in some pool of water which contained the common salts."

QUESTIONS

1. Why are sodium and silicon found in plants, when the latter have no use for them?
2. Where do animals get all the chemical elements necessary to build up their bodies?
3. Name the elements necessary for animal life.
4. What happens if one necessary element is lacking in a plant's or an animal's food?
5. What is the only uncombined element used by animals in their living processes? What do they use it for?
6. How would you prove whether chlorine is a necessary element for the corn plant?
7. Why do scientists believe that life originated in the sea?

LABORATORY EXPERIMENTS

51. To Show the Presence of Carbon in Plant Tissues.—Place some cornmeal, chopped hay, or other plant material in a porcelain dish and heat with the burner. The charring indicates carbon. What other substances have been proved to contain carbon? (See experiment 22.)

52. To Show the Presence of Some Mineral Elements in Plant Tissues.— (a) Ignite some chopped hay in a porcelain dish to a white ash, using two burners at once, if necessary. When cool, add water to the ash, filter a little of it, and test the filtrate for chlorides. (See experiment 50.)

(b) Add a few c.c. of dilute hydrochloric acid to the remainder of the ash in order to dissolve more of it, filter and test separate portions of the filtrates for phosphates and sulfates. To test for phosphates, make 10 c.c. of it alkaline with ammonia, then slightly acid with nitric acid. Heat to about 65° C. and add a few c.c. of ammonium molybdate solution.*

(c) Test for calcium by making one portion slightly ammoniacal and

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* The ammonium molybdate solution is made as follows: Dissolve 50 g. of molybdic acid in a mixture of 72 c.c. of concentrated ammonium hydroxide (specific gravity 0.90), and 135 c.c. of water; slowly and with constant stirring pour this solution into a mixture of 244 c.c. of concentrated nitric acid (specific gravity 1.42) and 574 c.c. of water. Keep the mixture in a warm place for several days. Carefully pour off the clear liquid from any sediment and keep in glass-stoppered bottles.

then adding a few c.c. of potassium or ammonium oxalate solution. The precipitate is calcium oxalate:

$$K_2C_2O_4 \;+\; CaCl_2 \;=\; CaC_2O_4 \;+\; 2\,KCl$$
$$\text{potassium} \qquad\qquad\qquad \text{calcium}$$
$$\text{oxalate} \qquad\qquad\qquad\quad \text{oxalate}$$
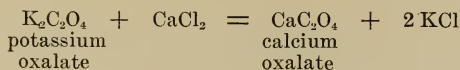
(d) If a portion of the ash will not go into solution it indicates silica.

(e) Name all the chemical elements that have been proved present in plant material, including the two in experiment 9. What other elements are probably present (Chapter XI)? These elements cannot easily be detected by tests, hence the tests have not been included in these experiments.

53. To Show What Mineral Elements Are Needed by Growing Plants.—Wash 4 or 5 pounds of the cleanest sand obtainable by stirring it in a pail of water, letting it settle, then pouring off the water. Repeat three times, then dry the sand.

Make up the following solutions:

I		II		III	
Water	1000 c.c.	Water	1000 c.c.	Water	1000 c.c.
KNO_3	1.0 g.	$NaNO_3$	1.0 g.	KNO_3	1.0 g.
NaCl	1.0 g.	NaCl	1.0 g.	KCl	1.0 g.
$MgSO_4$	0.5 g.	$MgSO_4$	0.5 g.	$MgSO_4$	0.5 g.
$CaH_4(PO_4)_2$	2.0 g.	$CaH_4(PO_4)_2$	2.0 g.	$CaH_4(PO_4)_2$	2.0 g.
$FeSO_4$	0.3 g.	$FeSO_4$	0.3 g.	$FeSO_4$	0.3 g.

IV		V		VI	
Water	1000 c.c.	Water	1000 c.c.	Water	1000 c.c.
KNO_3	1.0 g.	KNO_3	1.0 g.	KNO_3	1.0 g.
NaCl	1.0 g.	NaCl	1.0 g.	NaCl	1.0 g.
Na_2SO_4	1.0 g.	$MgSO_4$	0.5 g.	$MgSO_4$	0.5 g.
$CaH_4(PO_4)_2$	2.0 g.	Na_2HPO_4	2.0 g.	$CaH_4(PO_4)_2$	2.0 g.
$FeSO_4$	0.3 g.	$FeSO_4$	0.3 g.		

VII		VIII		IX	
Water	1000 c.c.	Water	1000 c.c.	Water	1000 c.c.
K_2SO_4	1.0 g.	K_2SO_4	1.0 g.	KNO_3	1.0 g.
NaCl	1.0 g.	$NaNO_3$	1.0 g.	$NaNO_3$	1.0 g.
$MgSO_4$	0.5 g.	$MgSO_4$	0.5 g.	$MgCl_2$	0.5 g.
$CaH_4(PO_4)_2$	2.0 g.	$CaH_4(PO_4)_2$	2.0 g.	$CaH_4(PO_4)_2$	2.0 g.
$FeSO_4$	0.3 g.	$FeSO_4$	0.3 g.	$FeCl_3$	0.3 g.

X	
Water	1000 c.c.
KNO_3	1.0 g.
NaCl	1.0 g.
$MgSO_4$	0.5 g.
$CaCl_2$	2.0 g.
$FeSO_4$	0.3 g.

Place about 2 inches of the sand in the bottom of 10 pint Mason jars. Label these jars as follows: I, complete; II, no K.; III, no Na; IV, no Mg; V, no Ca; VI. no Fe; VII, no NO_3; VIII, no Cl; IX, no SO_4; X, no PO_4. The

making of the solutions and the preparation of the jars may be divided up among the students. Place 10 timothy, wheat, barley or rye seeds in the sand in each pot, moisten the sand with the appropriate nutrient solution, and watch the development of the plants. Keep the jars covered with glass until the leaves appear, then leave them uncovered for the sake of aëration. Keep the sand damp, but not soaked completely with water. Every seventh day water with the nutrient solution and the other days with distilled water. If the nutrient solution is used up before the completion of the experiment, do not make up more solution, but used distilled water, since the above quantities of salts are sufficient for this number of plants. Higher amounts injure them.

Make careful observations on the development of the plants, noting particularly their height, color, and general thriftiness. Keep a record of all observations, noting which pots show the best growth.

Why was the sand washed? Why is distilled water used? What elements are apparently necessary for plant growth? The lack of what element causes the yellowing of the leaves?

CHAPTER XII

THE SOIL

Origin of the Earth.—Astronomers tell us that in the beginning the earth was a white-hot mass of gas and liquid, more or less like the sun is at present. It probably consisted of the same eighty chemical elements that now constitute the earth; but the temperature was so exceedingly high that none of the elements were combined; they existed as free elements. But the mass gradually cooled; and as the temperature became lower, some of the elements could combine. Thus, silicon and oxygen combined to form quartz, hydrogen and oxygen combined to form water, carbon and oxygen to form carbon dioxide. One after another, simple compounds were formed. Then some of these compounds combined to form other compounds, including minerals. During these changes the heavier elements and compounds gradually worked towards the centre of the ball of gas, and the lighter ones towards the outside. As the cooling process continued, more and more minerals and other compounds were formed; finally the greater proportion of the gas had cooled to a liquid state; and as it cooled still further, the molecules of one mineral would tend to collect together and crystallize into a solid, just as salt or sugar will crystallize from their concentrated solutions. Finally, the whole outer shell of the fluid ball became a solid crust, consisting of a large number of minerals in lumps cemented together. To be sure, some of these masses of minerals were as big as mountains, but they were small in comparison to the whole earth. One element, nitrogen, did not combine to any extent with other elements, and

remained on the outside of the ball as the atmosphere.
Another element, oxygen, was so abundant that although
it had combined to form nearly half of the weight of the
solid ball, there was still a great deal left over, which
stayed in the atmosphere with the nitrogen. The crust
of the earth, although solid, was still very hot; hence the
oxygen and the hydrogen that had combined to form
water could not remain on the earth as liquid water, but
existed in the air as water vapor. Probably all the carbon
in the world was in the form of carbon dioxide, and prac-
tically all of this was in the air, as the lime in the earth's
crust was still too hot to retain the carbon dioxide as cal-
cium carbonate. (See the chemistry of limestone, p. 100.)
Thus the earth's atmosphere at that time was consider-
ably deeper than it now is, and besides the present
amounts of nitrogen and oxygen, there was an immense
quantity of water vapor and carbon dioxide.

Beginning of Life and of Soil.—When the earth had
cooled sufficiently to allow water to condense on it, the
wear-and-tear process of erosion commenced. The water
falling in torrential rains flowed from the higher spots to
the lower, dissolving some of the minerals, exerting chemi-
cal action on others, and exposing new layers of them to
the atmosphere. It carried the material to the lower and
hotter portions of the crust, and then, evaporating, left
the mineral residue behind. In this way the primitive
seas were formed, and the highlands or mountains were
worn down. The waters of these seas thus contained
all of the chemical elements that had existed in the min-
erals which the water could dissolve from the mountains
and rocks. The minerals which the water could dissolve
and carry to the sea contained the fifteen elements dis-
cussed in the last chapter; that is, those necessary for
life. Hence the supposition that living organisms had
their origin in the sea.

The earth continued to cool; water and air continued to exert their action on the rocks and minerals; these were broken into smaller and smaller pieces by colliding against each other in streams; and still other processes of disintegration continued, over long periods of time. Finally, parts of the earth's surface were covered with fine rock débris. This was the beginning of soil. It was not soil as we ordinarily think of it until there were plants growing in it. This probably took place when the heat and moisture conditions were so favorable that some of the organisms living in the seas could gradually accustom themselves to living in the soil. When they did so, there was not only life in the sea, but life in the soil on land as well. More and more of the earth was covered with soil, more complicated forms of plant and animal life began to appear, the earth became sufficiently cool so that more and more of its surface would support life, the mountains became smaller and the seas larger, there was less water vapor in the atmosphere, a great deal of carbon dioxide was absorbed by plants, stored up in their bodies, and then the carbon converted into coal by the decay of the plants. Thus the world as we now know it was evolved. The process took an inconceivably long time, and of course was far more complicated than we have outlined. Furthermore, changes are still taking place; the earth will continue to cool off, the frigid zones will move farther and farther towards the equator, and finally the crust of the earth will become too cold to support life.

The Soil-forming Minerals.—Since our present soil had its origin mainly in the minerals of the earth's crust, it is important that we know something of the nature of these minerals, since each mineral furnishes a particular group of elements to the soil.

As we have seen above, the original rock of the earth's crust is a mixture of many minerals. *A mineral is a*

definite chemical compound, and each mineral always has the same composition. Granite, for instance, is not a mineral, for it varies in composition; it is a mixture of minerals, which can be seen in individual crystals if the freshly-broken surface is carefully examined, especially with a lens. A few of the more important minerals are described below:

Quartz, SiO_2, we have mentioned before as being the most abundant compound in the earth. All sands and sandstones are composed of it; the modified sandstone called quartzite is practically all silicon dioxide. Chirt, flint, agate, opal, and amethyst are forms of this mineral. It contributes no plant-food elements to the soil.

Feldspars form a group of very important minerals. Next to quartz, they are the most abundant of all minerals. All feldspars contain aluminum and silicon, and at least one other metal, which metal determines the kind of feldspar a particular one is. Thus, sodium, potassium and calcium feldspar are common, the most important being the potassium. The pink and green particles in granite are feldspars.

Micas are soft minerals, characterized by their tendency to split into thin, transparent sheets: The "isinglass" of coal stoves is mica. It is very abundant in granite rocks, constituting the shining flakes in them. They are compounds of aluminum, silicon, oxygen, and potassium, although some of them contain iron and calcium.

Apatite is essentially a phosphate and a chloride of calcium combined, and is practically the only phosphorus-bearing mineral in our rocks and soils.

Kaolin is an aluminum silicate with no other metal in its composition. It is the basis of all clays. With water it forms a sticky, adherent mass, which when baked becomes very hard and brittle. This property is valuable

in the making of bricks and pottery, as we have seen; but it is otherwise in soils, as clayey soils are noted for their stickiness in wet weather, and their hardness in dry, hot weather. Since kaolin results from the decomposition of micas, feldspars, and other granite minerals, clay soils are usually rich in plant-food elements.

Calcium carbonate and magnesium carbonate and their importance in the earth have already been discussed. Although very abundant in rocks, and sometimes forming deep beds and constituting whole mountains, they are only moderately abundant in most soils, since they readily leach out in water containing carbon dioxide, and in this way give rise to most of the hardness of water. River and sea waters contain a great deal of calcium and magnesium bicarbonates.

Selenite is calcium sulfate, called also gypsum and "land plaster." It is the principal source of the sulfur of plants. It is more soluble in water than most other minerals, hence is found abundantly in drainage waters.

Iron pyrites and limonite are two common iron-bearing minerals in rocks. The former furnishes the yellow and green colors to most sand and clay beds. These minerals are the source of iron to plants.

Talc is a magnesium silicate. It is commonly known in mass as soapstone, although in small granules it is a constituent of many rocks.

There are a great many more minerals found in all soils, but the above are the most abundant. They are also the most important, since they furnish all the elements needed by plants and by animals with the exception of iodine, which occurs in only minute traces in some little-known minerals.

Methods of Soil Formation.—Thus, we see briefly that the soil was formed by the disintegration of the minerals of the earth. To this decayed rock material was added de-

cayed plant material as soon as plants became abundant in the primitive soil. Our soils to-day contain a considerable and very important amount of decayed plant remains, called *humus*. The actual processes of soil formation involved several factors, the chief of which we shall now discuss.

Water is probably the most important agent in soil formation. Rain seeps into the crevices of rocks, dissolves out the most soluble minerals and leaves the less soluble behind. It even acts chemically on some minerals. When it contains carbon dioxide in solution it is ·very active chemically. We have already seen how limestone is dissolved by carbonated water:

$$CaCO_3 + H_2O = CO_2 + Ca(HCO_3)_2$$
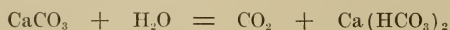
The calcium carbonate is very insoluble, the calcium bicarbonate very soluble. Many other minerals are acted upon similarly; as apatite, gypsum, feldspar, mica, and even quartz.

Running water wears away rocks by grinding one on another. It also carries away the finer particles and exposes new surfaces of rocks to the wearing action.

The *freezing* and *thawing* of water are powerful forces in splitting rocks. The water gets into crevices, and as it freezes it expands with great force and splits the rock. When the water thaws it flows into the crack a little farther, and on freezing continues the process. It is recognized by farmers that spring and fall weather of many alternate freezings and thawings is very beneficial to soils.

Frozen water in the form of *glaciers* has been the main agent in the formation of soils in most of our Northern States. In one of the later geological periods huge glaciers formed in Canada and slowly moved southward (Fig. 42). These masses of ice were hundreds of miles in extent,

and hundreds of feet in depth; they ground off whole hills, pulverized rocks, and dug deep gashes in the earth

Fig. 42.—Map of North America, showing the areas covered by the three great ice sheets of the Glacial Period. *A*, Labrador; *B*, Keewatin; and *C*, Cordilleran ice sheet. The masses of ice carried along enormous quantities of rock and soil, which were deposited when the ice melted. This material constitutes the present soil in the glaciated regions.

like the basins of the Great Lakes. And when they melted, they deposited their loads of rocks, gravel, and sand in piles. These gravel piles contribute most of the present

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hills in the glaciated areas of the Northern States and Canada. When the ice had melted and the climate became warmer, the surface of the earth in these regions was covered with a new layer of rock débris, on which the other soil-forming agents set to work and produced the wonderfully fertile soils which are there at present.

Winds hurl sand and particles against cliffs, and wear them off; blow the soil off of hills and expose new surfaces to the weather; and deposit fine silt particles over large areas to form the loess soils of our Western States.

The *roots of trees,* pushing their way into rock crevices and wedging them apart by the force of their growth, exert a great effect in the breaking up of rock into soil.

Animals, such as earthworms, ants, gophers, and moles, are important agents in moving soil about, aërating it, bringing subsoil to the surface, and aiding in the decay of plant remains.

Kinds of Soil.—It will be readily understood now how soils from various parts of the earth can vary so widely in composition, character, and fertility, depending first, upon the composition of the underlying bed rock, or, in the case of glacier, water, or wind-carried soils, upon the minerals of the region from which they came; and second, upon the methods of soil formation themselves. A soil might be low in phosphorus if there was little apatite in the rocks. A soil of granite origin is likely to be clayey, due to the decomposition of aluminum-bearing minerals. A soil deposited by wind will consist entirely of fine particles. A soil deposited from water, as the soils from old lake and river bottoms, is likely to be sandy. Soil in a glaciated region is likely to occur in patches of varied character and to be full of stones, both large and small. A soil in a limestone region may or may not be rich in lime, depending upon how extensive has been the leaching of carbonated water through it.

Soils are classified in many ways. Since the method of classification depends mainly upon the viewpoint, no arbitrary classification will be given here, but the discussion of soils will be given from various viewpoints. Thus, on the basis of position, we speak of *soils* and *subsoils;* from the standpoint of the rainfall of the region we have *arid* and *humid* soils; because of mechanical differences we have *sand, clay, loam, muck,* and *peat* soils; depending upon the origin and the method of formation we have *glacier, loess, sedentary, transported, limestone, granite,* and *lava* soils.

The most conspicuous difference between *soil* and *subsoil* is the color; this is due to the decayed plant matter or humus in the soil. There is usually less calcium carbonate, potassium, and iron compounds in the soil, as these are leached away and used by plants to a considerable extent. The subsoil is more fine grained and compact, as the finer particles of the soil are washed down into it. For this reason the subsoil is heavy and is not easily aërated, and hence is much less fertile than the soil.

Arid soils are not subjected to the leaching action of water to the extent that *humid* soils are; hence the amount of available plant food is usually greater in the arid soils. They were formed not so much by the action of air and water on the minerals as by the alternate freezing and thawing of water in the rock crevices. When arid soils are supplied with water, as by irrigation, they usually prove to be very fertile.

By *sand* is usually meant the coarser particles of broken rock, consisting mainly of silicon dioxide, but also of many other minerals in smaller amounts. The term "sandy soil" means nothing definite. By it is usually meant a soil of rather coarse, loose, open texture, containing little humus, holds water but poorly, and is rather infertile until organic matter has been added to it. *Silt*

consists of particles smaller than sand, and silt soils have a closer texture and hold water better than sandy soils. The loess soils of the Western plains consist mainly of silt deposited by winds. *Clay* is that portion of soil which consists of the very finest of particles; the particles of clay are so fine that they do not settle when suspended in water. Since clay consists of finely ground-up mineral materials of all kinds, it contains an abundance of plant food. It holds water tenaciously, forming a sticky, adhering mass, thus making a soil that is worked with difficulty. The term *loam* is the most abused of all the words in soil terminology and usually means the least. Any soil which will grow crops, is somewhat dark in color, and is worked fairly readily is a loam; and this is usually modified to "sandy loam," "clay loam," and other terms. *Peat* is partially-decayed plant remains, as we saw in the chapter on coal. It occurs in marshes where certain characteristic grasses and mosses thrive. It contains very little mineral material, although what there is becomes readily available to plants when the peat marsh is drained and aërated. It absorbs water like a sponge; hence it is used to pack around the roots of nursery stock, and one variety is used in place of absorbent cotton in surgical dressings. When peat is so thoroughly decayed that plant stems and leaves can no longer be distinguished in it, and it has become very black in color and has been mixed with mineral soil to some extent, it is called *muck*. Muck is the typical soil of marshes.

Soil types which derive their names from their method of origin have already been discussed sufficiently under the topics of the methods themselves, as *loess, glaciated,* and *water-deposited* or *sedimentary* soils. Our *granite* and *lava* soils are usually very rich in plant-food elements, because the appropriate minerals are abundant in these rocks. Calcium carbonate is usually low, however, as it is

leached out as rapidly as the soil is formed. We saw that limestone had its origin in the skeletons of sea-animals; these skeletons carried with them parts of the softer portions of the animals as well, and thus limestone usually contains an appreciable amount of phosphorus, potassium, iron, magnesium, and other elements. When the limestone disintegrates in soil formation, the calcium carbonate is the most soluble of all the constituents; hence the resultant limestone soil contains relatively little calcium carbonate and relatively much other minerals. Such soils are very fertile, although often they are benefited by applications of ground limestone. The majority of the soils of the Mississippi Valley are of limestone origin.

The Mechanics of Soil.—The capacity of a soil to hold water and to absorb water, the rapidity with which it becomes warm in the spring, and the ease with which it can be plowed and hoed and kept in a pulverized condition, are spoken of as the *mechanical properties* of the soil. These depend almost entirely upon the proportion of sand, clay, and humus in the soil. Sand holds water very poorly; therefore it is quickly drained in the spring, and it as quickly becomes too dry later in the summer to support plant growth. The particles do not adhere to any degree, hence sandy soils are always loose, easily worked and well aërated. Clay absorbs water very greedily; it also retains it just as greedily, and thus prevents plants from absorbing it. Thus a sandy soil containing 10 per cent of water will still keep plants from wilting, where a clayey soil with 15 per cent of water will not. If clay soils are worked when they are wet, they "puddle," and form sticky lumps that, when dry, are hard, brick-like clods. This can be remedied by incorporating sufficient sand with the soil, or, more easily, by adding finely ground limestone, which "flocculates" the clay and reduces the tendency to puddle.

Warm and Cold Soils.—The rapidity with which a soil becomes warmed by the sun is governed almost entirely by the amount of water in the soil. This is because water has a very high specific heat; that is, it requires more heat to raise the temperature of a pound of water one degree than that of any soil constituent. Therefore wet, poorly drained, clay and muck soils are always the last to become warm enough in the spring to promote plant growth. Drainage is the only cure for such soils.

The *humus* of soils also largely governs the water-holding capacity, as it absorbs water even more readily than does clay. The following table shows this in actual figures:

	Pounds of water in 1 cu. ft. of soil
Soil high in sand	28
Soil high in clay	42
Soil high in humus	51

This explains why the addition of manure and other humus-forming material greatly benefits sandy soils. Since black materials always absorb more light and heat waves when exposed to the sun's rays than do materials that are not black, humus serves to produce warmer soils by absorbing more of the sun's rays. Humus is more important in soils chemically than mechanically, however, and it will be taken up from this standpoint later on.

Chemical Composition of the Soil.—We have seen what the origin of soils is; what minerals furnish the chemical elements necessary for plant growth; what effect the source of a soil and the method of its formation may have upon the amount of these elements in the soil. We have yet to discuss the actual amounts of these elements present, and how the growth of crops and the chemical activities in the soil may change these actual amounts.

The accompanying table gives the approximate composition of soils, both in percentages and in the amounts

of the various constituents in an acre six inches, assuming that the latter weighs 2,000,000 pounds.

TABLE X

Average Composition of Soil

	Percentage	Amount in acre 6 inches
Silicon dioxide........................	50–80	1,700,000 lbs.
Aluminum............................	8–14	200,000 lbs.
Iron................................	1–6	80,000 lbs.
Calcium..............................	0.5 –2.0	30,000 lbs.
Magnesium...........................	0.5 –1.5	20,000 lbs.
Sodium..............................	0.1 –1.3	16,000 lbs.
Potassium............................	1.0 –3.0	40,000 lbs.
Sulfur...............................	0.02–0.2	2,000 lbs.
Chlorine.............................	0.20–1.2	16,000 lbs.
Phosphorus...........................	0.03–0.15	1,600 lbs.
Nitrogen.............................	0.02–0.10	1,000 lbs.
Humus..............................	0.02–4.0	40,000 lbs.

Total Plant Food.—It will be seen that although the percentage of some of the plant-food elements is small, the total amount in an acre is very large. Also, it is apparent that sulfur, phosphorus, and nitrogen are the least plentiful of these elements. When the above figures are compared to those in the table (p. 182) which gives the amounts of the various elements contained in an acre of various crops, it will be seen that even the scarcest elements in the soil are still sufficient to supply crops for a great many years. Thus, a thirty-bushel crop of wheat contains about nine pounds of phosphorus; 1600 pounds of this element in an acre six inches of soil would then support about 175 such crops. Similarly, there is enough of the other scarcer elements to support ordinary crops for several hundred years. We know, however, that continuously cropped soils become unproductive long before that time. In other words, soils become infertile while there is still an abundance of plant-food elements in them.

The reason for this lies in two facts which we have already considered. The first is, that plants require all their food in solution. The second is, that the decomposition of soil and rock minerals into water-soluble forms is a slow, gradual process.

Available Plant Food.—Now, the analyses of soils in the table on page 167 give the *total* amount of those elements in the soil, irrespective of whether they are in a soluble or insoluble form. Since the greater proportion of each of them is still in the insoluble mineral form, such analyses really give us very little idea of the fertility of a soil. Thus in soil studies we have to deal with two classes of soil constituents, those that are available to plants, and those that are unavailable.

It is often found that the total phosphorus or potassium in a soil is very large, sufficient for years of bumper crops, and yet that soil will not produce those crops. The application of fifty pounds of readily-available phosphorus or potassium fertilizer, however, works wonders in increasing the fertility.

Chemical Analysis of Soils.—It might seem that all the chemist would have to do would be to dissolve out with water the water-soluble material and this would be the portion available to plants. Unfortunately, however, soil behaves so differently towards water in this form from what it does towards the soil water and the roots of plants, that such a water extract gives very little information as to the fertility of the soil, or as to the fertilizer requirements. In fact, *one thing that agricultural chemistry has failed to give us so far is a method by which we can judge by chemical analysis the value of a soil.* To be sure, an analysis might suggest that more humus might be beneficial, or that lime is needed to correct acidity, or that some one element is suspiciously low; but it could not give us grounds for prophesying what crop would grow

best, or what fertilizing element would be most effective.

Factors in Soil Fertility.—In soil fertility, then, we have to deal with a number of factors. The soil must absorb large quantities of water, and yet it must not become water-logged and soggy, and it must give up the water readily to plants. It must not only contain sufficient available plant food for all crop requirements at the time being, but it must also contain an abundant supply of unavailable food, which can be constantly and steadily reduced to an available form as fast as needed. It is this last process of rendering the insoluble minerals into available form, that the farmer has under his control to a considerable extent. The chemical changes that take place in the soil and the methods of controlling them will be discussed in the next chapter.

QUESTIONS

1. What has been the most important agent in the wearing down of mountains and in the formation of soil? In what different ways does it act?
2. Where did life probably originate?
3. When could plants begin to live in soil?
4. Make a list of the most important soil minerals, indicating what chemical elements necessary for plant growth are found in each.
5. State the various soil-forming agents in the order of their importance.
6. How can soils be classified?
7. What is sand? Silt? Humus? Subsoil? Limestone? Peat? Loam?
8. Explain how the moisture-holding capacity of a soil depends upon the amount of sand, clay, and humus in it.
9. What is meant by total plant food and by available plant food in soils?
10. What does a chemical analysis of a soil indicate?
11. What are the most important scarce elements in soils?

LABORATORY EXPERIMENTS

54. To Show the Presence of Various Mineral Elements in Soil.—
(a) Leach about 100 g. of sandy loam with hot water acidified with nitric acid, filter through double filter paper until the filtrate is as clear as possible, then test for chlorides.

(b) Boil 50 g. of soil with 100 c.c. of 20 per cent hydrochloric acid for 10 or 15 minutes, cool, filter, wash the soil on the filter somewhat, then use the filtrate for testing as follows:

(c) Add a few drops $BaCl_2$ solution to a few c.c. of it. This tests for what?

(d) Make 10 c.c. of it faintly ammoniacal. What is the precipitate? Add a few drops of a solution of potassium oxalate. The granular precipitate is calcium oxalate.

(e) If the solution is yellow, iron is present, since $FeCl_3$ is yellow in solution.

(f) Make 10 c.c. of the solution alkaline with ammonia, then slightly acid with nitric acid; warm to 65° C. and add a few c.c. of ammonium molybdate solution. What does this test indicate? Why couldn't the extract of soil in (b) be used for testing for chlorides? A great many other elements are found in soil, but their tests are not sufficiently simple to be given here.

55. To Study the Humus in Soil.—Shake about 100 g. of black soil for one minute in a 200 c.c. cylinder containing 1 per cent hydrochloric acid. Let the soil settle, then pour off the acid. Repeat, then repeat twice with water. Then add 100 c.c. of 4 per cent ammonia and shake. The brownish or blackish colored material that dissolves is the humus combined with ammonia. Acidify some of the humus extract with hydrochloric acid. Explain the result. Evaporate the rest of the solution to dryness in an evaporating dish. Hold a flame under the dish. The charring proves that humus contains what?

56. To Compare the Water-holding Capacity of Different Kinds of Soil.—Obtain some dry sand, loam, clay, and peat. Weigh out 20 g. of each in shallow dishes, then pour water from a graduated cylinder on to them until they are thoroughly saturated with water, and the water is just beginning to run away from the substances. Record the c.c. of water absorbed by the same amounts of these materials. How can the moisture-holding capacity of a soil be improved?

57. To Show the Effect of a Mulch Upon the Evaporation of Water From a Soil.—Fill two tin or glass containers of the same size to the depth of two inches with loam soil. Pack the soil down somewhat firmly, then wet thoroughly. Stir the surface of the soil in one dish until it is thoroughly tilled. Then weigh each dish, and set aside in a warm place. Weigh each day thereafter, noting any effect that the mulching may have on the rate of water loss. What is the fundamental principle of dry farming?

58. To Show the Effect of Color on the Temperature of Soil.—Fill two boxes about one foot square with soil. Cover the surface of one with ground charcoal, bone-black, or some other black powder, and the other with ground gypsum, chalk, or other white powder. Insert the bulb of a thermometer in the soil in each box and expose the boxes side by side to the direct rays of the sun. Record the temperature in each box every 15 minutes for two hours.

CHEMICAL CHANGES IN THE SOIL

FROM the previous discussions we have seen that the soil is not a fixed material, with unchanging composition, but that it is the seat of ceaseless activities and changes. Rocks and minerals are constantly being decomposed into soil particles; some of the soluble material from these is being washed away; plant remains are decaying in the soil and forming new compounds; rain brings down foreign substances from the atmosphere; manure and fertilizers are being added. What chemical changes do these things actually bring about? Which changes are good and which are bad, and how can they be controlled?

Decomposition of Minerals.—As we learned in the last chapter, only a small proportion of the total mineral nutrients of a soil is available at any one time for crop use. It is very necessary, then, that fresh supplies of these nutrients be made available continuously. One of the most active agents in decomposing minerals is water containing dissolved carbon dioxide. We have seen how this affects limestone (p. 102). The familiar fertilizer, rock phosphate, is practically insoluble in water; in the presence of carbon dioxide, however, it is converted into a soluble phosphate. Feldspars and micas have their potassium set free under the influence of carbonated water. It is therefore important that soil water contain plenty of carbon dioxide; and this is furnished by the humus.

Formation of Humus.—Humus is the decayed or decaying organic matter of the soil; it consists of the remains of plant and animal bodies which are being used as food by bacteria, fungi, and molds, and by them being broken

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down into a complex mixture of substances which have a black color. Usually, the blacker the soil the more humus it contains.

Mechanical Effects cf Humus.—Humus is a gummy, sticky material. Since it thus serves to cement soil particles together, the addition of humus to sandy soils makes them more compact. As we have mentioned before, humus absorbs water very greedily. It therefore aids the water-holding capacity of sandy soils. On clayey soils, however, it apparently has an opposite effect; and the addition of manure and other humus-forming materials to clays is well known to be very effective in ''loosening up'' their texture and making them more tillable.

Chemical Properties of Humus.—The above functions of humus are physical or mechanical. Its chemical properties are probably far more important even than these, for the following reasons: (1) The humus contains all of the nitrogen of the soil; (2) it furnishes all the carbon dioxide for the soil water; (3) it furnishes the food for the micro-organisms which help to decompose the soil minerals; (4) it is the source of the acidity of soils. Each of these points is so important that it is worthy of more lengthy consideration.

The nitrogen of soils is not found in the minerals, for there are practically no nitrogen-containing minerals in our soils. The residues of the plant and animal bodies that are returned to the soil contain the nitrogen for the succeeding crops. This nitrogen is in a multitude of different compounds, because all the compounds of nitrogen in the roots and stems of plants, in the bodies of insects and earthworms, in the packing house by-products used as fertilizers, in the urine and fæces of farm animals, are present in the soil; and on these compounds the bacteria set to work to produce humus. In feeding on these materials, the bacteria produce still other compounds, among

the most important of which are carbon dioxide, acids, ammonia, and nitrates. The carbon dioxide is important in increasing the solvent power of soil water on minerals. The acids create the harmful effects noticed in "sour" soils. The ammonia and nitrates are compounds of nitrogen, and as such are of first importance in feeding the growing plants.

Nitrification.—Now, plants are very peculiar in their nitrogen nutrition. *It is not sufficient that the compound of nitrogen simply be soluble in water; the nitrogen must be in the form of nitrates, or, with some plants, in the form of ammonia.* This means that if a soil is to supply plenty of nitrogen to growing crops, the bacteria of the soil must be constantly busy converting the various kinds of humus nitrogen into nitrates. As soon as they fail to keep up the supply of nitrates, the crops begin to suffer nitrogen starvation. *This process of forming nitrates out of the nitrogen compounds of plant and animal residues is called nitrification.* The first step is usually ammonification. The ammonia is then converted into nitrates. The ease with which a nitrogen-containing material can be nitrified determines to a large degree its value as a fertilizer. For example, dried blood and manure are excellent nitrogenous fertilizers, because the soil bacteria can very quickly nitrify them; but ground leather and hair scrap, although high in nitrogen, are almost worthless as fertilizers, because it is years before they are completely nitrified. The above facts also serve to explain why saltpeter (sodium nitrate) is such a quick-acting fertilizer; it needs no modification in the soil, but is immediately utilized by the plants.

Assisting Nitrification.—Thus one of the most important chemical changes in soil is that brought about by the nitrifying bacteria. The farmer can assist this process by keeping the soil well supplied with nitrogenous organic

matter for the bacteria to act upon. Barnyard manure is one of the best and cheapest of such fertilizers. The roots and stubble of leguminous crops, such as alfalfa and clover, are also excellent sources of nitrogen. Some of the commoner nitrogenous commercial fertilizers will be discussed in the next chapter. Upon the physical condition of the soil also depends largely the rate of nitrification. It is a process requiring oxygen; therefore good aëration is necessary. This can be secured by tillage and by keeping clay soil coagulated by means of limestone (p. 165). The latter also neutralizes any acidity, which is harmful to the nitrification process.

Denitrification.—The opposite of the above process, *denitrification*, also takes place in the soils. And the conditions which favor denitrification are the opposite of those which favor nitrification. A cold, wet, sticky, non-aërated soil is very likely to have denitrification taking place. This process involves loss of nitrogen; the nitrates lose their oxygen, and the nitrogen is liberated as free nitrogen, and hence is of no further use to plants. By working for conditions which favor nitrification, a farmer is thus also preventing denitrification.

Nitrogen Fixation.—It was indicated above that the principal source of nitrogen in soils is the nitrogenous compounds contained in the plant and animal bodies that decay in the soil, and in the manure and fertilizers that are added to the soil. There is, however, another very important source of nitrogen, and that is the uncombined nitrogen of the air. Although our ordinary crop plants demand nitrate nitrogen, and can make no use whatever of the nitrogen all around them in the air, there are certain kinds of bacteria which can take this atmospheric nitrogen and use it for building up their bodies. These are called nitrogen-fixing bacteria. There are two classes of them. The members of one class live in the humus of the soil, and

being microscopic in size, are never seen and would not be
known to exist there, were it not for the increase in the
percentage of nitrogen found by chemical analysis in soils
richly supplied with them. The other class of nitrogen-
fixing bacteria live on the roots of clover, alfalfa, beans,
peas, and other plants of the legume group (Fig. 43).

Fig. 43.—Effect of high nitrogen and of low nitrogen on legumes. The cowpeas on the
right had a plentiful nitrogen supply; hence they developed no nodules of nitrogen-fixing
bacteria on their roots. The cowpeas on the left had a scant nitrogen supply; they developed
nodules, thus enabling them to draw upon the air for a great deal of their nitrogen. It is
evident that for a legume to bring about the greatest increase in soil nitrogen the soil must
be rather low in nitrogen to start with. (From Bulletin 230, Wis. Agric. Exp. Station.)

They form nodules or lumps on the roots sometimes as
large as peas, and are often prominent features of the root
systems of these plants. The bacteria and the plant have
a well-ordered system of coöperation. The plant gives
the bacteria a place to live and furnishes them with certain
mineral nutrients; in turn the bacteria absorb nitrogen

from the air, and manufacture it into compounds which they turn over to the plant for its use. In this way the plant obtains the nitrogen for its growth without drawing upon the supply of nitrogen in the soil. Then, when the stubble of this crop, as well as the manure resulting from the feeding of the crop to stock, is plowed into the soil and allowed to decay, *the total nitrogen in the soil has been increased by the amount taken from the free nitrogen of the air.*

Importance of Legume Bacteria.—Let it be perfectly understood that these *legume bacteria* carry on an ex-

Plants

Soil   *Animals*

N in air ———→ *Legumes* ————→ *Animals*
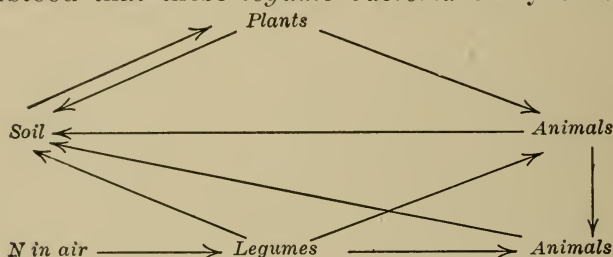
FIG. 44.—Diagram showing the nitrogen cycles. The transformation of the free nitrogen of the air into compounds of nitrogen in legumes, and from the legumes to the nitrogen compounds in animals, in the soil, and finally in other plants, can be followed by means of the arrows.

ceedingly important chemical reaction in the soil; and that it is a chemical reaction which is very much under the control of the farmer. The plant can get along without the bacteria on the roots if there are plenty of nitrates in the soil; but the soil does not gain in nitrogen thereby. By planting a legume crop, seeing to it that the proper bacteria are in the soil for infecting the roots, liming the soil if it is too acid for the proper growth of the bacteria, and then returning to the soil as much as possible of the manure from the crop, the farmer can very materially increase the store of nitrogen in his soil. Figure 44 shows graphically the interrelation of the nitrogen in air, soil, plants, and animals.

Inoculation for Legumes.—Experience has shown that the same variety of bacterium does not inhabit all the legume plants; hence a field that will develop nodules on

Fig. 45.—Last step in preparing a seed-bed. The whole operation kills weeds, thoroughly mixes the soil with the manure and fertilizer that have been added, grinds up the débris from the previous crop, which is to make manure, and produces a dust mulch which is so necessary for the prevention of loss of water by evaporation.

red clover will not necessarily develop on alfalfa or soy beans. It depends upon whether nodule-bearing alfalfa or soy beans have been grown on that field before. Therefore inoculation is often resorted to when the bacteria

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for a given kind of legume are lacking in a field. The inoculation can be carried out either by obtaining some soil from a field that is known to contain the desired organisms, or by sending to the Department of Agriculture at Washington, or to some private company, for cultures. The seed is inoculated in either case; the specific directions are obtained from any of the state agricultural experiment stations, or from the package in the case of the culture method.

Soil Acidity.—Bacterial action on humus gives rise to various organic acids, which, if allowed to accumulate, produce the condition known as ''sour'' or acid soil. This seriously inhibits the growth of many crops, interferes with nitrification, and prevents the development of nitrogen-fixing bacteria on legume roots. Therefore, this acidity must be remedied by neutralizing with finely-ground limestone. Thus limestone, besides neutralizing soil acidity, has other beneficial effects, such as furnishing calcium for plant food, assisting in nitrification by forming calcium nitrate from the nitric acid produced, and flocculating or ''loosening up'' clay soils. Hence the spreading of ground limestone has become almost as common a farm operation as the spreading of manure.

Cultivation and Chemical Changes.—Lastly, we must consider how mechanical treatment of soil can assist chemical changes. Plowing and cultivating keep soil in a pulverized, porous condition (Fig. 45). This allows free exchange of gases between the soil and the atmosphere. And since oxygen is necessary in bringing about nitrification and in preventing denitrification, and since atmospheric nitrogen is necessary for the nitrogen-fixing bacteria, it will be readily seen why thorough cultivation and aëration of the soil are necessary in the proper control of soil activities.

Thus we see that the soil is the scene of many well-

defined chemical changes. When the nature of these reactions, the conditions under which they take place, and their desirability or undesirability in the soil are understood, the management of the soil can be pursued more intelligently and more profitably.

QUESTIONS

1. What purpose does carbon dioxide in the soil serve?
2. What is the most important chemical element in humus? Why?
3. What mechanical properties of humus are important in soils?
4. Why is humus important in the soil chemically?
5. What is nitrification? Of what use is it in soils?
6. Explain fully how the nitrogen content of a soil can be increased by growing alfalfa.
7. How can denitrification be prevented?
8. Why is an acid soil undesirable? How can it be corrected?
9. In what ways does cultivation assist in the chemical changes in the soil?

LABORATORY EXPERIMENTS

59. To Show That Lime Flocculates Clay.—Mix about 20 g. of clay soil with water in a mortar to a thin cream. Pour this into a 200 c.c. cylinder of water, mix, and allow to stand. The clay remains suspended in the water while the other constituents settle out. Divide the suspension of clay into two portions in glass cylinders. To one cylinder add 0.5 g. of fresh quicklime, and mix. Let the two cylinders stand for several days, shaking occasionally, and noting any change that takes place in the clay. When this flocculation takes place in clay soil, the latter loses much of its stickiness.

60. To Show the Presence of Nitrogen in Soil.—(a) Grind together in a mortar a few grams each of sodium hydroxide with a similar quantity of flour, ground hay, or other plant or animal material. Place the mixture in a test tube, and across the opening of the test tube lay a damp piece of red litmus paper. Carefully heat the contents of the tube. The alkali decomposes the nitrogenous compounds of the flour, converting the nitrogen into ammonia. The ammonia affects the litmus paper and it may even be smelled.

(b) Repeat the above test for nitrogen, using about 10 g. of rich loam soil and 1 g. of NaOH. What is the source of the nitrogenous substances of the soil?

CHAPTER XIV

MANURES AND FERTILIZERS

Need of Fertilizers.—We saw in the last chapter that for a soil to be fertile and to remain fertile, the following conditions must hold true:

1. The rainfall, temperature, and sunshine must be suitable.

2. The soil must contain sufficient mineral plant-food elements in an available form, and the supply of this available food must be maintained year after year.

3. The mechanical condition must be good, for proper tillage, moisture-holding capacity, etc.

4. The soil must not contain injurious chemical substances, nor microörganisms which will cause plant diseases.

It is obvious that the first is out of human control. The others, thanks to extensive and careful studies in soil science by the state and government experiment stations in this country and in Europe, are under human control, and this control is almost entirely chemical. Hence, we shall see that one of the greatest contributions that chemistry has made to the practice of modern agriculture is in the maintenance of soil fertility and the rendering of infertile soils fertile.

Manure is the general term applied to anything added to a soil to improve its fertility. It does not necessarily add plant food elements; it may correct the physical condition of the soil, or destroy a harmful constituent. This broad use of the term manure is an old one; at the present time we apply different names to the various classes of manures, as follows:

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1. *Barnyard manure* is the excrement of farm animals plus the litter that is usually mixed with it.

2. *Green manure* is a crop or portion of a crop which is plowed under while still green, to add humus to the soil.

3. *Commercial fertilizers* are materials containing plant food elements, which are bought and sold on the markets.

4. *Soil amendments* are substances added to soils to destroy harmful ingredients or to improve in some way the fertility, without adding plant food.

1. **Barnyard Manure.**—The use of animal excrements to enrich the soil is as old as the practice of agriculture. In fact, the better and more permanent the agriculture practiced by a given people, the more do we find that they use barnyard manure. Compare the history of agriculture in our New England States with that in China, for example. In these States, where farming has been carried on continuously for a hundred years or so without very careful attention to the return of manure to the soil, the soil is now "worked out," and so unproductive that in many places the farms have been abandoned. In China, a dense population is still supported on land that has been continuously cultivated for many hundreds, perhaps even thousands, of years; and it is simply due to the fact that they have carefully conserved all animal and even human excrement and returned it to the soil.

Loss of Plant Nutrients from the Farm.—Let us see why the careful utilization of barnyard manure is so intimately connected with the maintenance of soil fertility; let us especially see what the chemistry behind it is.

We grow an acre of wheat. A good crop, according to the accompanying table, will remove from the soil about 50 pounds of nitrogen, 9 of phosphorus, and 6 of sulfur. These amounts were taken from the *available* supply in the soil; hence chemical action must convert an equal

amount of the *unavailable* supply of these elements into available, and do so year after year, if that soil is to continue producing good crops. If the grain is sold off the farm, and the straw is stacked up and burned, that much nitrogen, phosphorus, etc., is lost to that farm entirely. If, however, the grain and straw are fed, the excrement of the animals will contain almost the whole amount of these elements, and if the excrement is returned to the land without wastage, the soil will not suffer any appreciable loss by the growth and harvesting of this crop, and the next crop will not be so dependent upon the unavailable supply in the soil.

TABLE XI

Amount of Plant Food Removed from the Soil by Various Crops, Expressed in Pounds Per Acre

	Dry weight of crops	N	P	S	K	Ca	Mg
	lbs.	lbs.	lbs.	lbs.	lbs.	lbs.	lbs.
Alfalfa hay................	6000	132	11.2	16.7	118	101	21.6
Red clover hay............	3700	97	11.2	6.1	67.2	63.4	14.9
Turnips (roots)............	3000	61	9.7	23.0	92.1	18.0	3.5
Potatoes (tubers)..........	3300	46	9.4	1.1	64.5	2.4	3.7
Corn (grain) 30 bu........	1500	28	4.3	2.5	5.4	0.3	2.0
Corn (stalks)..............	1800	14	3.6	2.2	24.7	7.5	3.3
Total for corn.........	3300	42	7.9	4.7	30.1	7.8	5.3
Wheat (grain) 20 bu........	1100	21	4.3	1.7	5.0	0.4	1.4
Wheat (straw).............	2400	14	2.5	3.4	15.8	5.2	2.0
Total for wheat.....	3500	35	6.8	5.1	20.8	5.6	3.4

Permanent fertility of a farm, therefore, is largely dependent upon, first, the keeping of stock on that farm, and second, the conservation of the manure from that stock (Fig. 46).

Composition of Manure.—What is a ton of manure worth to the soil? This depends, it may be said, mainly upon

two factors: the composition of the fresh manure, and the
composition of the manure when it is plowed under the
soil; for from the moment the excrement leaves the ani-
mal, it is subject to a continuous series of chemical
changes. Some of these changes are good, some are very
bad; and the extent to which these changes are controlled

FIG. 46.—Effect of manure on corn grown on depleted soil. The unmanured plot (on the
left) yielded 36 bushels per acre, the manured 62 bushels. Subsequent crops for two or
three years will still show the benefits of this application of manure. (From Bul. 222, Purdue
Agric. Exp. Station.)

determines the measure of conservation practiced on a
given farm.

Effect of Kind of Animal.—In Table XII are presented
the analyses of the fresh manure (solid and liquid com-
bined) of the various farm animals.

The value was computed by assuming the values of
35 cents, 9 cents, and 20 cents per pound of nitrogen,
phosphoric acid and potash, respectively. These are
conservative figures, based on the 1920 prices of these
ingredients in commercial fertilizers. It is seen in the
table that the water content of the manures is the deciding

factor in their value; that the amount of water differs with the different animals; and that the nitrogen is the most prominent fertilizing constituent. Compared to commercial fertilizers, barnyard manure has a low value per ton because it is over three-fourths water. However, if we take $5.50 as the average value of the above manures at the present time, this is over twice the value assigned to manure in 1914. This is due to the enormous increase in the price of potash since the opening of the World War. The above values of manure include only the nitrogen, phosphorus, and potash, based on what they cost in commercial fertilizers; it takes no account of the organic matter in the manure, which forms humus in the soil, nor of the loosening effect of manure on heavy soil. A dollars-and-cents value cannot be assigned to these things, of course.

TABLE XII

Composition and Value of a Ton of the Various Farm Manures

	Pounds of				
	Water	Nitrogen, N	Phosphoric acid, P_2O_5	Potash, K_2O	Value *
Cow.................	1540	8.8	3.2	8.0	4.96
Pig..................	1460	9.0	3.8	12.0	5.89
Horse...............	1400	11.6	5.6	10.6	6.68
Sheep...............	1280	16.6	4.6	13.4	8.90
Hen.................	1100	20.0	16.0	8.0	10.04

* Value computed on basis of 1920 prices; 35 cents a pound for nitrogen, 9 cents for phosphorus, and 20 cents for potash.

Effect of Feed on Manure.—Now, the figures in the accompanying table show only how the kind of animal affects the composition of the manure. Two other factors, however, must be considered. One is *the kind of feed,* the other is *the kind and amount of bedding* used. As regards the feed, the following general rule applies: The daily excrement (urine plus fæces) of animals contains the

same amount of fertilizing elements as their feed. This applies strictly, of course, only to those animals which are neither gaining nor losing in weight. Important exceptions to the rule are the following: Young growing animals, which retain part of these elements to build up new tissue; milch cows, which secrete these elements in the milk to a considerable extent; pregnant animals, which require these elements for the growing fœtus; and hens laying eggs. We can say, in general, for all classes and conditions of farm animals, that four-fifths of the nitrogen, phosphorus, and potash of the feed gets into the excrement. Therefore, if feeds are chosen which are richer in these materials, the resultant manures will also be richer. The accompanying table gives the value of the fertilizing constituents in some of the common feeding stuffs.

TABLE XIII

The Value of the Nitrogen, Phosphorus, and Potash in One Ton of Various Feeds and Litters

	Pounds of			
	Nitrogen, N	Phosphoric acid, P_2O_5	Potash, K_2O	Value *
	lbs.	*lbs.*	*lbs.*	
Wheat middlings.................	52.6	19.0	12.6	$22.61
Corn (grain).....................	36.4	14.0	8.0	15.60
Cottonseed-meal.................	13.5	57.6	17.4	13.49
Corn silage.....................	5.6	2.2	7.4	3.53
Whole milk.....................	10.0	3.0	3.0	4.37
Wheat straw....................	11.8	2.4	10.2	6.38
Oat straw......................	12.4	4.0	24.8	9.66
Sawdust........................	4.0	6.0	14.0	4.74

* Value computed on basis of 1920 prices; 35 cents a pound for nitrogen, 9 cents a pound for phosphoric acid, and 20 cents a pound for potash.

Effect of Bedding on Manure.—As regards the effect of bedding on the composition of manure, it is obvious that the richer the bedding is in fertilizing elements and the more of it that is used, the greater will be the yearly manurial output of a farm. Table XIII contains data concern-

ing the value of wheat straw, oat straw, and sawdust as litter. It must be remembered, in comparing these three, that the sawdust decays much more slowly than the others.

Fermentation in Manure.—The fermentation processes in manure are brought about by the myriads of microorganisms that are always present in the intestinal tracts of animals. Two general types of fermentation are recognized: (1) Hot, or dry, fermentation, also called firefanging, which takes place in dry manures, such as that of the horse; and (2) cold, or wet, fermentation, such as takes place in a compacted pile of wet manure, like that of the cow.

It is well recognized that the first kind is undesirable; a great deal of the organic matter is destroyed during the fire-fanging, and considerable loss of ammonia occurs. The latter is often recognized by its odor, especially in the urine of horses. The ammonia comes from the urea, according to the following equation:

$$\underset{\text{urea}}{(NH_2)_2CO} + \underset{\text{water}}{2\,H_2O} = \underset{\text{ammonium carbonate}}{(NH_4)_2CO_3}$$

The ammonium carbonate decomposes as follows, giving off ammonia gas:

$$\underset{\substack{\text{ammonium}\\\text{carbonate}}}{(NH_4)_2CO_3} = \underset{\substack{\text{ammonia}\\\text{gas}}}{2\,NH_3} + H_2O + CO_2$$
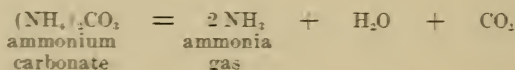
The loss of this valuable nitrogenous constituent can be prevented in the stable by scattering a little gypsum, or land plaster, $CaSO_4$, in the litter:

$$(NH_4)_2CO_3 + CaSO_4 = \underset{\substack{\text{calcium}\\\text{carbonate}}}{CaCO_3} + \underset{\substack{\text{ammonium}\\\text{sulfate}}}{(NH_4)_2SO_4}$$
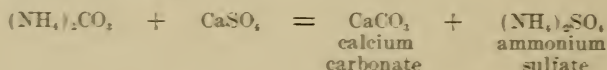
The ammonium sulfate is not volatile.

Rotting of Manure.—The kind of manurial fermentation to be encouraged is the wet type. The bacteria feed upon

the substances of the manure; they decompose the fibrous residues of the undigested food; disintegrate the materials of the bedding; and change the various compounds of nitrogen, phosphorus, and other elements into forms which are available to plants. If air has access to the fermenting material, considerable heat may develop. The air should therefore be excluded, by keeping the piles closely compacted and rather wet. The deep stall method of conserving manure, the mixing of the wet and the dry manures, the saving of all the liquid excrement, the use of water to keep the piles sufficiently wet, are all effective methods of securing well-rotted piles of manure with a minimum of waste.

Losses of Manure.—Precaution to secure well-rotted manure is all for naught, unless the opportunities for wastage and loss are carefully guarded against. Thus, rain pelting over the pile washes away the soluble portion of the manure, which is the portion most quickly available to plants; the running off of the liquid excrement is a direct loss of readily available plant food; and the dumping of manure in little piles in the field results in "spotted" effects in the next crop, and the failure to secure the most value from the fertilizer. All in all, the best possible practice in the handling of manure is to spread it on the fields fresh from the stables daily (Fig. 47).

2. **Green Manures.**—These are crops planted for the purpose of being plowed into the soil while they are still green, in order to furnish easily-decayed humus-forming material to the soil. Usually they are legumes, although non-legumes, such as rye, are sometimes used. The economy in using a legume is apparent, since it not only adds humus to the soil, but also adds the nitrogen which it absorbs from the air during its growth. Whole crops are plowed under for green manures only in those systems of

farming where the amount of stock kept is not sufficient to supply the needed manure. For it is an obvious waste to plow under feed crops which could better be fed and the manure returned to the soil. It was found at the Rothamsted Experiment Station in England that the roots and stubble alone of an acre of clover contain about 50 pounds of nitrogen.

3. **Commercial Fertilizers.**—These ordinarily are de-

Fig. 47.—Manure carrier and manure spreader, two machines which facilitate the handling of manure. If manure can be hauled to the fields daily, all loss by leaching and by harmful fermentation is avoided.

signed to supply but three plant-food elements—nitrogen, phosphorus, and potash. The raw materials are certain basic substances which usually contain but one of the fertilizing elements. Very often two or more of these are mixed and put on the market as ''mixed'' or ''complete'' fertilizers.

Nitrogen Fertilizers.—The nitrogenous fertilizers most commonly used are Chile nitrate (sodium nitrate, or Chile saltpeter), ammonium sulfate, calcium nitrate, dried blood,

and fish scrap. The Chile nitrate is mined in Chile, where it occurs in enormous beds, and from which over a million tons a year of the fertilizer are obtained. It contains about 15 per cent of nitrogen. It is very soluble in water, and is, of course, already nitrified and ready for use by plants. Hence, its greatest use is in forcing greenhouse and truck crops, where it is applied in small amounts several times during the season in order to prevent leaching out by rains.

Ammonium sulfate is a product of the gas works (p. 50). It is very soluble in water and is readily nitrified. It contains from 20 to 23 per cent of nitrogen.

Calcium nitrate, as we have seen (p. 128), is a manufactured product. It can be produced cheaply only where there is cheap electric power, which means abundant water power. Its development in this country is full of promise; when we begin to utilize our many water-power sites for generating current, calcium nitrate is bound to become an inexpensive but very valuable fertilizer. It contains about 13 per cent of nitrogen.

Dried blood is a packing-house by-product containing from 10 to 14 per cent of nitrogen. It readily undergoes nitrification; hence it is largely used for forcing truck crops and greenhouse plants.

Fish scrap is usually the residue left after extracting the oil from menhaden, although the residue from canning factories is made into fertilizer to a certain extent. It contains from 8 to 10 per cent of nitrogen, besides 4 to 5 per cent of phosphoric acid. This fertilizer nitrifies nearly as readily as dried blood.

Phosphorus Fertilizers.—The phosphorus fertilizers most commonly used are rock phosphate, acid phosphate, ground bones, and basic slag. The *rock phosphate* goes by the name Florida rock, Tennessee rock, etc., depending upon its source. It is insoluble in water, although when

ground very fine it is rendered soluble by the action of carbonated water (p. 171). Hence the advisable practice is to add it to soil in connection with manure. It is essentially tri-calcium phosphate, $Ca_3(PO_4)_2$, and contains from 25 to 30 per cent of phosphoric acid.

When rock phosphate is treated with sulfuric acid, *mono-calcium phosphate,* also called *acid phosphate* and *super-phosphate,* is formed:

$$Ca_3(PO_4)_2 \;+\; 2\,H_2SO_4 \;=\; CaH_4(PO_4)_2 \;+\; CaSO_4$$
$$\text{rock phosphate} \qquad\qquad \text{acid phosphate} \qquad \text{calcium sulfate}$$
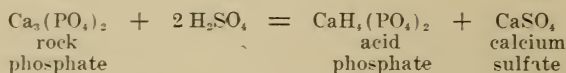
This acid phosphate is soluble in water, and hence is a more quickly-acting fertilizer than the rock phosphate. It contains from 14 to 16 per cent phosphoric acid. The calcium sulfate formed in the above reaction is not removed from the phosphate; hence the fertilizer as it appears on the market contains both substances. In many instances the beneficial effects from the use of this fertilizer are due as much to the sulfate portion as to the phosphate.

Ground bones are either "raw" or "steamed." The latter form is preferable in some ways, since the steaming extracts all fat and thus enables the material to be decomposed more readily in the soil. This material contains from 25 to 30 per cent of phosphoric acid in the form of tri-calcium phosphate, besides some nitrogen.

Basic slag is a material arising from the removal of phosphorus and silicon in the making of iron and steel. It contains from 15 to 20 per cent of phosphoric acid in an insoluble form. However, if it is very finely ground, it readily becomes soluble in the soil by the action of salts and of carbon dioxide.

Potassium Fertilizers.—Up to the opening of the World War the greater portion of the world's potash fertilizers

came from the Stassfurt mines of Germany. These mines are immense deposits of water-soluble potassium and magnesium salts, from which several different fertilizing materials are drawn. The muriate, or chloride, KCl; the sulfate, K_2SO_4; and kainit, $K_2SO_4.MgSO_4.MgCl_2.6H_2O$, are the principal ones. The first two are the most concentrated, averaging nearly 50 per cent of potash. Kainit is a crude salt mixture, usually containing not over 13 per cent potash.

American Sources of Potash.—When the importation of Stassfurt potash ceased, the fertilizer chemists of America had to make a careful inquiry into all the potassium possibilities on this side of the Atlantic. After several years of effort, there are now being placed on the market potash salts from a variety of sources. One of them is the giant sea-weeds of the Pacific coast. These are raked out of the water, fermented in huge tanks, and then the potash, together with several other by-products, is recovered from the liquor as potassium chloride. Another source of potassium is a by-product in the manufacture of Portland cement, where feldspars high in potash are used as raw material. At Searles Lake, California, there are immense deposits of salts of various kinds, partly crystalline and partly in a strong brine. This brine is pumped out into tanks, evaporated, freed from the bulk of its sodium chloride by crystallization, and then evaporated to dryness. This product contains about 75 per cent potassium chloride. There are 30,000,000 tons of this salt available in this lake.

Wood ashes are a not insignificant potash fertilizer. Unleached hardwood ashes contain as high as 8 per cent of potash in the form of carbonate.

4. **Soil Amendments.**—Under this head, lime in various forms is about the only substance used in this country. As was stated in the chapter on soils, ground limestone

is often added to soils to neutralize excess acidity, since
most crops thrive best in a neutral soil. The limestone
also tends to make sticky clay soils more loose and easier
to work, while on light, sandy soils it apparently has oppo-
site effects. Quicklime, CaO, is often recommended in
place of ground limestone. Since it is caustic, it cannot be
applied when any vegetation is likely to be injured by it.
And since, as we learned in our study of the chemistry of
calcium compounds, calcium oxide is soon converted into

Fig. 48.—Efficiency of limestone of different degrees of fineness. The more finely it is
ground, the more intimately it can be mixed with the soil, and the more quickly will it neu-
tralize acidity and improve the crop yield. The quicklime (on the left) slacks to a very
fine powder. This is all that is gained by using quicklime instead of ground limestone.
(From Bul. 152, Pa. Agric. Exp. Station.)

the hydroxide by moisture, and the hydroxide is very
quickly converted to the carbonate by the carbon dioxide
of the air, it is obvious that very little can be gained
in speed of action or in total action by the use of quick-
lime, provided that the limestone used be of good quality
and very finely ground (Fig. 48). In neutralizing power,
a ton of quicklime is equivalent to about two tons
of limestone.

Fertilizer Terms.—It is very unfortunate that a loose
system of terms has come into use in the fertilizer trade,
in respect to the different forms of nitrogen, phosphorus.

and potash. It may be well to point out the terms for each of these elements and their real significance.

The stated analysis of nitrogen fertilizers should consist of the per cent of N, together with a statement as to the form of nitrogen, whether nitrate, ammonia, or organic (as in dried blood and fish scrap). However, the nitrogen figure is often computed to the ammonia (NH_3) basis, and given as the percentage of ammonia present, whereas there may be no ammonia nitrogen in it. The ammonia figure is 17/14 of the nitrogen figure; hence it gives a misleading idea as to the composition of the fertilizer. Again, the same figure may be stated in two different forms: as, "per cent of nitrogen, 8; ammonia equivalent, 9.7"; and many purchasers believe that they are getting both, or a total of 17.7 per cent of fertilizing constituents.

The term "potash" in fertilizer nomenclature almost invariably refers to K_2O. This potassium oxide never occurs in any fertilizer; but it has become customary to figure the amount of K present in terms of K_2O, and then to speak of it as potash instead of potassium. The actual compound of potassium present should be stated.

In the phosphate fertilizers, the phosphorus is computed as P_2O_5 instead of P, and is called phosphoric acid. In reality it is phosphorus pentoxide, or phosphorus anhydride. It is well to know that all of these terms mean the same. The kind of phosphorus compound involved should be stated, so that the purchasers may know something of its availability to plants.

Mixed Fertilizers.—Any one of the commercial fertilizers discussed above may be purchased singly, and used alone or in conjunction with barnyard manure. Most of the fertilizers sold, however, are "mixed" or "complete" fertilizers, and contain all three fertilizing elements. Their composition is stated in the form of three figures, as 3—8—5, which signifies 3 per cent of nitrogen, 8 per

13

cent of phosphoric acid, and 5 per cent of potash. The farmer can buy the individual fertilizers and mix them himself. This has the advantage of being somewhat cheaper and of giving a fertilizer of known composition. The disadvantages are, that the single fertilizers cannot always be bought on the market, great care is necessary to secure thorough mixing, and an intimate knowledge of the character of the materials to be mixed is necessary, else combinations may be used which will react with each other chemically and result in loss of ammonia or in the formation of insoluble phosphorus and potassium compounds.

Choice of Fertilizers.—One of the first questions that confronts a farmer in his endeavor to produce maximum yields of crops is, will fertilizers pay? If so, what should be used? As was pointed out in the chapter on soils, a chemical analysis gives but a general idea of the nature of the fertilizer requirements. A far surer way is to establish fertilizer plots on the most typical soil type of a farm, and to try out the various fertilizer elements singly and in combination, on one or more main crops. Usually it is advisable to apply ground limestone to all plots, as often the possible effect of fertilizer will fail to appear because of the reaction of the soil. Control plots receiving no fertilizer must of course be kept; and the crops from each plot should be accurately determined by weight, so that any difference in yield can be recognized on plots as small as one-tenth or one-twentieth acre. This method requires time and labor; but it gives accurate information as to the efficiency of fertilizers on that farm for those particular crops. A convenient system of plots, together with the amount and kind of fertilizer to add to each, is given in the accompanying table. Each plot is one rod wide and eight rods long, containing one-twentieth of an acre.

TABLE XIV

Plan for Fertilizer Test Plots

Plot No.	Kind of fertilizing element	Plant food applied
1	None	None.
2	N	Dried blood, 30 lbs.
3	P	Acid phosphate, 15 lbs.
4	K	Potassium sulfate, 15 lbs.
5	None	None.
6	N, P, and K	Blood, phosphate, and sulfate.
7	N and P	Blood and phosphate.
8	K and P	Sulfate and phosphate.
9	N and K	Blood and sulfate.
10	None	None.

Specific Effect of Fertilizers.—In a general way, nitrogenous fertilizers favor the development of stem and leaves; phosphorus develops the fruit and pods; and potassium, the roots and stems. Thus, nitrogen fertilizers are particularly effective on truck crops such as lettuce, onions, cabbage, and rhubarb, and on hay crops and meadows. Phosphorus fertilizers are usually applied to the cereals and to corn, while potassium is usually very effective on tobacco, potatoes, and root crops (Fig. 49). However, so much depends upon the particular soil involved that the above statements are only generalities. Deficiency in nitrogen is usually indicated by a light green or even yellowish color of the leaves, and underdeveloped leaf and stem parts. Delayed maturity, and undersized seeds, usually mean too little available phosphorus.

Systems of Fertilization.—Several well-defined systems for applying fertilizers have been suggested, and practiced to a certain extent. The *Ville* system depends upon the assumption that crops can be divided into three groups, each group being particularly in need of an excess of one fertilizing element. Thus nitrogen should be the dominant element in fertilizers for wheat, rye, oats, bar-
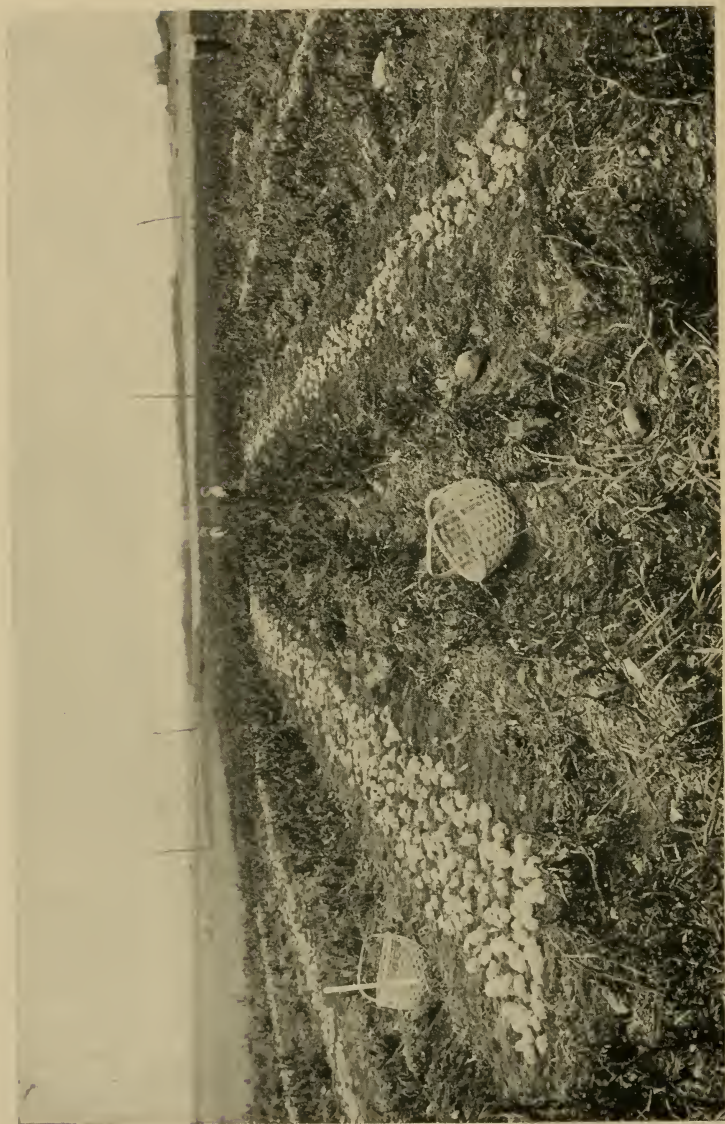
FIG 49.—Effect of commercial fertilizer on the yield of potatoes. On the left, two fertilized rows; yield, 364 bushels per acre. On the right, two unfertilized rows; yield, 226 bushels per acre. Potatoes usually respond to potash fertilizers.

ley, grass, and beets; phosphorus for turnips, corn, sorghum, and sugar-cane; and potash for peas, beans, clover, flax, potatoes, and tobacco. This system is of considerable help in choosing fertilizers for intensive operations when there is a reasonably good supply of all other elements of fertility.

The *Wagner* system assumes that nitrogen is the only critical element in soil fertility, because it is found in soils in very small amounts, is easily leached out, and is expensive to buy in fertilizers. Hence, phosphorus and potash are supplied in reasonable amounts to all crops, and then nitrogen supplied at frequent intervals in readily available forms, such as ammonium sulfate or sodium nitrate. This system implies relatively cheap phosphorus and potassium fertilizers, so that they can be present at all times in more than necessary quantities. Like the Ville system, this one would pay only when intensive agriculture is practiced.

The *money crop* system involves applying all the fertilizer to that crop in the rotation which is sold from the farm, in order that the maximum effect of the outlay for fertilizer may be had as a cash return. The unused portion of the fertilizer goes to feed the succeeding crops. This in general is the system followed in this country.

QUESTIONS

1. What are the main conditions on which soil fertility depends? Which of these are under the control of the farmer?
2. Name the four classes of substances that are added to soils to improve their fertility, and state the purpose of each class.
3. Why is barnyard manure the best all-round material for increasing soil fertility? Name three different things that it does for the soil.
4. Why is a ton of sheep manure more valuable than a ton of cow manure?
5. What is the most important plant food element in manure? Why (three reasons)?
6. Explain why stock must be kept on a farm if the soil fertility is to be maintained year after year.
7. Explain how the kind of feed and the kind and age of the animal affect the composition of the manure.
8. What is the effect of fire-fanging?

9. Why does manure ferment and rot so readily?
10. Explain fully why legumes are the best green manures.
11. Make a table giving the principal kinds of nitrogenous, phosphatic, and potassic fertilizers, the source of each, and the percentages of plant food elements.
12. State three ways in which limestone may improve the fertility of a soil.
13. By means of the various tables in this chapter, compute the amount of nitrogen, phosphorus, and potash in the manure produced in one year on a farm stocked with 6 horses, 4 full-grown hogs, 16 pigs, and 100 hens. Assume that (1) all of the manure is left on the farm and is hauled to the field without loss; (2) wheat straw is used for bedding, allowing 5 pounds per day for each horse, and 2 pounds for each hog; (3) each hen produces 140 pounds of manure a year.
14. What is a single fertilizer? A complete fertilizer? When is it economy to use each?
15. How is nitrogen starvation shown in the appearance of a plant?
16. Discuss briefly the principal systems followed in the use of commercial fertilizers.
17. Why cannot fertilizers take the place of barnyard manure entirely?

LABORATORY EXPERIMENTS

61. To Show the Presence of Nitrogen in Manure and Fertilizers.— Use the test for nitrogen described in experiment 60, on dry manure, dried blood, ammonium sulfate, sodium nitrate, rock phosphate, and " muriate of potash " (potassium chloride). Why did not the saltpeter give the test? Which of the above fertilizers are nitrogenous?

62. To Show the Solubility of Rock Phosphate.—Suspend 2 g. of finely ground rock phosphate in 20 c.c. of water. Divide into two test tubes. Bubble CO_2 through one test tube, and allow the other to stand meanwhile, with occasional shaking. Then test for the amount of dissolved material, as was done in experiment 13. Test each filtrate for phosphorus. How does an abundance of humus in soil aid in the solution of rock phosphate?

63. To Prepare Acid Phosphate.—Thoroughly mix 10 g. of bone ash or ground rock phosphate with 6 c.c. of concentrated sulfuric acid and let stand on a piece of wood for two or three days. How has the mass changed? What does it consist of? Grind it up and test its solubility, comparing it with that of the rock phosphate in experiment 62.

$$Ca_3(PO_4)_2 \quad + \quad 2H_2SO_4 \quad = \quad CaH_4(PO_4)_2 \quad + \quad 2CaSO_4$$
tricalcium   monocalcium
phosphate   phosphate
(bone ash)   (acid phosphate)
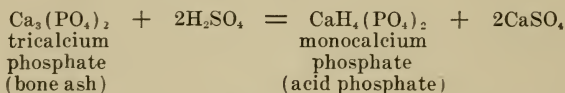
64. To Study Wet and Dry Fermentation of Manure.—Fill a one-pint Mason jar half full of loosely packed fresh horse manure. Half fill another jar with the manure packed down tightly and moistened with additional water. From the cover of each jar suspend a moist strip of red litmus paper. Stand in a warm place, noting any difference in temperature between the two jars, and any difference in the tint of the litmus paper. Discuss the difference in the chemical changes taking place in the two jars.

CHAPTER XV

THE PLANT BODY

Importance of Plants.—We have already mentioned the fact that probably the first forms of life to appear on the earth were simple plants. It is reasonable to think that this was the case, since how could animals live, unless plants were there for them to feed on? At the present time plant life is the basis for all animal life, including even the complex life of human civilization. Let us enumerate some of the ways in which we are dependent on plants in our everyday life.

1. Coal, which furnishes the greater part of the energy of our factories, locomotives, and boats, and supplies our buildings with heat, is the product of the decay of luxuriant plant life that once flourished on the earth.

2. It is probable that our petroleum and natural gas also had an origin in plant products.

3. Our clothing comes mostly from the cotton and the flax plant. The fabrics which have an animal origin, such as wool and silk, come from plants indirectly, since the sheep and the silkworm feed on plants.

4. Our food is really all from plant sources, for while part of it is the meat and milk and fat of animal bodies, the animals had to eat great quantities of plant material to produce these products.

5. All of our paper is plant material, mostly from wood, straw, and flax.

6. The lumber of our buildings, cars, and ships is from plants.

In view of these facts, then, it will be readily agreed that the primary business of agriculture is the production

of crops, and that the production of animals is only sec-
ondary. We have discussed the chemistry of the soil and
the air, which are the sources of the raw materials for
plants. We have next to study the chemical make-up of
plants and of animals, and the processes by which the
plant, animal, and mineral substances are converted into
our various manufactured products.

Raw Materials in Soil and Air.—We have seen in previous
chapters that the plant is composed of some twelve or
fourteen chemical elements combined in a large number of
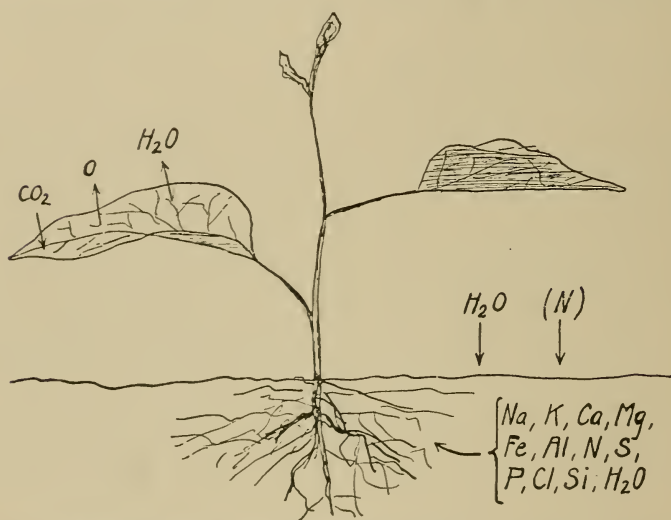
FIG. 50.—Diagram of a green plant, showing its relations to the supplies of raw materials
of the air and soil which it uses in building up its tissues.

different compounds; that the plant draws upon the soil
and the air for its supply of these elements; that the
nutrient minerals of the soil are carried into the plant by
means of water and that hence they must be soluble in
water; that the carbon dioxide of the air furnishes all the
carbon of the plant; and that this carbon dioxide enters

the plant through the minute pores in the leaves, called stomates.

Processes in the Plant.—Figure 50 is a simple diagram to show the processes in the plant factory. In the chapter on carbon we found that the carbon dioxide is taken into the tissues of the leaf, the carbon combined with water to form starch, and the oxygen again given off from the leaf as a waste product. This chemical change is indicated in the diagram by CO_2 entering the leaf and O issuing from it.

Importance of Water.—Water is shown both entering the roots and evaporating from the leaves. This indicates that more of it enters the plant than is required in the tissues, and that hence the excess must be eliminated. This passage of water through the plant is called *transpiration*. As a matter of fact, an enormous quantity of water passes up through a plant and out into the air again. For every plant there appears to be a minimum amount of water which is required during the growing season in order for the plant to attain a vigorous mature development. This *water requirement* of some of our cultivated crops is shown in the following table:

TABLE XV

Water Requirement of Some Common Plants

Pounds of water transpired for each pound of dry matter produced

Western wheat grass	1,070	Potato	640
Western ragweed	950	Oats	600
Flax	900	Barley	530
Alfalfa	830	Wheat	510
Red clover	800	Corn	370
Pumpkin	790	Sorghum	320
Cotton	640	Millet	310

These figures mean that for every pound of dry matter produced by alfalfa plants, 830 pounds of water had to be absorbed from the soil. The total water transpired from an acre of crops is thus very large. It will be

noticed that sorghum, one of our most drought-resisting
plants, has one of the lowest water requirements; and
that the opposite is true of the legumes, which are very
dependent on an adequate water supply. Marsh plants
have the highest, and desert plants have the lowest, rate
of transpiration.

It has been demonstrated many times that more water
is taken from an acre of soil when it is covered with
growing plants than when it lies fallow. In other words,
transpiration is much more rapid than evaporation from
bare soil. This explains the practice in some dry-farming
regions of allowing the land to lie fallow every other year;
the rain water is better retained in the soil when there is
no crop to use it up.

The Plant Factory.—We must look upon the plant as a
chemical factory. And it is a most efficient and remark-
able factory, too. As we have seen, the primary chemical
reaction that is brought about by green plants is the
formation of starch from carbon dioxide and water. The
plant does this quietly and unpretentiously, without any
outward manifestation except a slow growth of the various
organs of the plant. And man, in spite of all his progress
in manufacturing the substances of nature, in spite of his
powerful chemicals, electric currents, high temperature,
and machinery, has never yet succeeded in forming starch
from carbon dioxide and water. Therefore we must culti-
vate crops which will do this for us.

And why is the manufacture of starch so important?
In the first place, when a plant has once made some starch,
it can convert it into the various other substances that go
to make up its body, such as the fats, sugars, acids, and
coloring matters. In the second place, starch is an im-
portant human food. In the third place, we can feed it
to animals and they will convert it into the various animal
food products. Therefore, we must appreciate the plant

factory which makes such large quantities of this raw
material, starch, out of such plentiful substances as carbon
dioxide and water.

Photosynthesis.—We have learned before that the above
process in plants is called *photosynthesis,* which means
"manufacturing by means of light." Light, however, is
not the only requirement. The following conditions must
obtain before photosynthesis can take place:

1. The plant must be living. Dead cells, even though

Fig. 51.—A common fungus. The fungi (commonly called toadstools and mushrooms)
form a group of plants which do not contain chlorophyll; hence they cannot manufacture
their own food out of water and the carbon dioxide of the air. Therefore they live on the
living or dead tissues of other plants, and get their nourishment from them. When they
live on living plants, we speak of them as plant diseases. Thus fungi cause such diseases
as the rusts, the mildews, the smuts, late blight of potatoes, chestnut blight, brown rot of
peaches, and cabbage wilt. Large species of fungi like the one in the illustration are most
excellent human foods, although some are poisonous. (From Bul. 175, U. S. Dept. of
Agriculture.)

they are supplied with all the following conditions, can-
not manufacture starch.

2. The plant must be a green plant; that is, its cells and
tissues must contain the green coloring matter called
chlorophyll. Thus tubers and roots, the bark of trees,
ripe seeds, and mushrooms and yeast, are incapable
of photosynthesis (Fig. 51).

3. The plant must have a plentiful water supply, for
several reasons: First, because water is needed to com-
bine with carbon dioxide to form starch; second, because
when a plant is suffering for water, the pores in the leaves

become closed, and carbon dioxide from the air cannot get into the leaf; and third, because all the mineral material is brought into the plant in solution in water.

4. There must be a proper temperature. Photosynthesis, as well as many other chemical reactions in plants, practically ceases at 35 or 40 degrees F. On the other hand, too high a temperature is detrimental.

5. The plant must have sunlight. Some plants can thrive in indirect light, in shaded places, but practically all of our crop plants require direct sunlight. At night, photosynthesis ceases. Crops can mature in our northern latitudes principally because the daylight hours are longer, thus enabling the plants to obtain the necessary amount of sunlight in a shorter growing season. Light is a form of energy just as heat and electricity are. Plants, in some way not very well understood, can utilize sunlight as a source of energy for carrying out the manufacturing process of photosynthesis.

If any one of the above five conditions is lacking, photosynthesis cannot take place.

Chemical Compounds in Plants.—We must now take up the systematic study of the various substances that go to make up the bodies of plants. The accompanying table is a scheme to show a classification of these substances, together with the chemical elements of which they are composed. For convenience they are first grouped as water and dry matter. Of the dry matter one portion is combustible, and another portion is not, being left behind as ash. It has been pointed out before that the most abundant element in the dry matter of plants is carbon; in the above outline by far the greater part of the carbon appears in the combustible portion, although a little appears in the ash in the form of carbonates. For convenience the organic substances are divided into two groups, those containing nitrogen and those without.

TABLE XVI

An Outline of the Compounds Found in Plants

Water H and O			
Dry matter	Organic or combustible matter C, H, O, N, S	Non-nitrogenous C, H and O	1. Carbohydrates 2. Fats 3. Plant acids 4. Volatile oils 5. Coloring matter
		Nitrogenous C, H, O, N, S	6. Proteins 7. Amides 8. Alkaloids
	Mineral matter or ash Salts of Na, K, Ca, Mg, Fe, Al, P, S, C, Si, and Cl		

In the green plant, *water* occupies a very large proportion of the total weight. Lettuce and celery often consist of from 92 to 94 per cent water; potatoes and roots, 85 to 90 per cent; the driest hay and straw contain at least 8 per cent of moisture; and even wheat flour has 10 per cent. The water is not only the medium for bringing food into the plant from the soil, but it distends the individual plant cells, and thus keeps the plant organs firm and rigid. The familiar wilting of succulent plants shows the effect of insufficient water on the turgidity of leaves.

1. The *carbohydrates* are a very widely distributed group of compounds found in both plants and animals. They are classified into sub-groups as follows:

a. Pentoses, of the formula $C_5H_{10}O_5$
b. Hexoses, of the formula $C_6H_{12}O_6$
c. Di-hexoses of the formula $C_{12}H_{22}O_{11}$
d. Poly-hexoses of the formula $(C_6H_{10}O_5)x$

It will be noticed from the formulas that each carbohydrate is composed of carbon, together with hydrogen and oxygen in the proportion of 2 to 1, as in water. Thus, the pentoses consist of 5 atoms of carbon combined with 5 molecules of water. Hence the name carbohydrate.

These constituents can be demonstrated by placing some sugar in a test tube and slowly heating it. Water is given off, and condenses on the sides of the tube, while the residue gets blacker and blacker, due to the carbon liberated. Pure carbon for making artificial diamonds is obtained by roasting pure cane sugar until it is completely decomposed and all the water is driven off. The water is, of course, chemically combined with the carbon; man cannot bring about this combination by chemical means, but the plant can by the process of photosynthesis, discussed in previous pages.

a. The *pentoses* are the simplest natural sugars. They are found in the gums which exude from cherry, apple, and spruce trees; in gum arabic; in straw and hay to the extent of 20 to 30 per cent; in wheat-bran to the extent of 24 per cent; and in all fruits and vegetables in small amounts. They are fairly well digested and assimilated by animals, but it is doubtful whether they are of much use to the human system. The pentoses contain five carbon atoms in the molecule, hence the prefix *pent.*

b. The *hexoses,* or monosaccharides, contain six carbon atoms combined with six molecules of water, as their formulas indicate. The commonest members of this group of carbohydrates are the three sugars, *levulose, dextrose,* and *galactose.* These sugars all have the same chemical formula, $C_6H_{12}O_6$; but they are entirely different from each other in sweetness and in other characteristics. The explanation of so many different substances having the same formula is, that although there are the same number of atoms of carbon, hydrogen, and oxygen in each, these atoms are arranged differently, and hence give different structures to the molecules. It is the same as if three dog-houses were built each with the same number of bricks and boards, but each having a different shape and appearance.

Levulose, also called fructose and fruit sugar, is found in practically all fruits and is especially abundant in honey. It is the sweetest of all sugars, which fact explains the intense sweetness of honey.

Dextrose, also called glucose and grape sugar, is found in at least small amounts in practically all plants; it is

FIG. 52.—A field of sugar cane ready to harvest. The cane is ripe in this country after about 11 or 12 months' growth. The leaves are stripped from the stalks by hand, the stalks cut with a heavy knife, and then hauled to the mill for pressing. (Copyrighted, Underwood and Underwood, N. Y.)

especially abundant in fruits and in sweet corn, and is the only sugar found in the blood and muscles of animals. It is about two-thirds as sweet as cane sugar.

Galactose is not found as such in nature, but is one of the constituents of milk sugar, as will be seen under that heading.

c. The *dihexoses* or disaccharides, as the name implies, contain twice the number of carbon atoms that the hexoses contain. Here again the same formula, $C_{12}H_{22}O_{11}$, applies to three different sugars—sucrose, lactose, and maltose—each differing from the other in sweetness, solubility, etc.

Sucrose, also called saccharose, is our most familiar and common sugar. Granulated sugar, from either beet or cane, is pure sucrose. Maple sugar is also sucrose, but contains certain impurities which give it its delightful flavor. ''Brown sugar'' is cane sugar that has not been completely purified. The sugar crystals that settle out from sorghum syrup are sucrose. Besides occurring to the extent of about 15 to 18 per cent in sugar cane (Fig. 52) and in sugar beets (Figs. 53 and 54) and about 12 per cent in sorghum, it is found in lesser amounts in sweet corn, sweet potatoes, watermelons, and in practically all fruits.

Fig. 53.—Diagram of sugar-beet, showing the sugar content of the different sections of the root. In the beginning of the beet sugar industry, the average sugar content was about 9 per cent; plant breeders have now developed them to yield an average of over 16 per cent. (From Circular 34, Utah Agric. Exp. Station.)
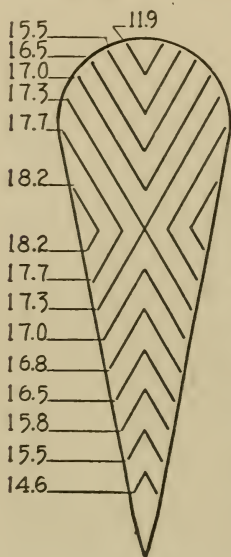
The production and manufacture of granulated sugar constitute one of our greatest industries. In normal times there are consumed in the United States alone something over 85 pounds of sugar for each man, woman, and child, or about 7 pounds per capita per month. When, during the World War, this was reduced to 2 pounds we were so accustomed to unlimited amounts of sugar, candy, ice cream, and other confections, that we found it difficult to adjust ourselves to the change, and to adopt other more available sugars for our source of sweets.

In the *manufacture of cane sugar*, the juice is expressed between a series of heavy rollers, and then clarified and purified to free it from coloring matter and from other non-sugar substances. Charcoal and sulfur dioxide are important agents for these purposes (pages 64 and 136). The clarified juice is then evaporated to a syrup. When this syrup reaches the proper consistency, the

Fig. 54.—Sugar-beet field and beet sugar factory. The sugar-beet is a strong rival of sugar-cane in producing the world's supply of sugar, although the beet has been cultivated for this purpose for comparatively few years. In the beet sugar industry the work of the chemist begins in the field with the selection of the best roots for seed stock, and ends in the factory, when the greatest possible amount of crystallized sugar has been recovered from the roots and the pulp and molasses have been converted into stock feeds. (Courtesy of the director of the Utah Agric. Exp. Station.)

sucrose crystallizes out. The mass of crystals and molasses is placed in a centrifugal machine. This consists of a perforated drum revolving rapidly inside of another drum. The force of the whirling tub throws the molasses out through the perforations, leaving the crystals of sucrose behind. When these are washed, dried, and sacked, the granulated sugar is ready for the market. A second crop of crystals is usually obtained from the

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molasses and then the final molasses is sold for human food or made into stock feeds.

The *manufacture of beet sugar* is similar to the above, except that the beets have to be washed, cut into thin slices, and then loaded into huge tanks, where the sugar is extracted by soaking in warm water. This water solution of the sugar, containing also non-sugar impurities, is put through a process similar to that of the cane juice. The extracted pulp as well as the molasses is used for stock feed.

Lactose, called also milk sugar, is the carbohydrate found in milk to the extent of about 5 to 6 per cent. It is not found in the plant world. It is less sweet than sucrose.

Maltose, or malt sugar, is the sugar found in malted or sprouted grain. It occurs in only small amounts before germination; but during the latter process the starch is converted into maltose.

d. The polyhexoses, or polysaccharides, are so called because their molecules are made up of a large number of molecules of hexoses combined with the elimination of a molecule of water. Thus the formula $(C_6H_{10}O_5)x$ means that an unknown, or x number of molecules of hexoses, each one minus a molecule of water, have combined to form a polyhexose. All the mono- and di-hexoses are sugars, and thus have a sweet taste and form crystals. The polyhexoses have no particular taste and do not crystallize. The commonest polysaccharides are starch, dextrin, glycogen, and cellulose.

Starch occurs very abundantly in the plant kingdom; it is the commonest form in which plants store up a reserve supply of carbohydrate for the use of the seedling plant the next year. Thus, practically all seeds, tubers, and roots contain starch. It is found in these organs in the form of minute grains of characteristic size and shape. These grains are insoluble in cold water; but on boiling

they swell up and form the familiar starch paste. Raw starch is digested with great difficulty by the human system, but readily by animals. Cooked starch, however, either as pure starch or as it occurs in grains and tubers, is readily digested and constitutes our most important carbohydrate food.

Dextrin is formed by heating starch above the boiling-point of water. It is the constituent on the crust of bread and in toast that is slightly sweet and becomes sticky when moistened with water. It is prepared in large quantities from starch for use as a mucilage, as on the backs of stamps and envelopes.

Glycogen is sometimes called animal starch, since it is very similar to starch and is found in the liver and muscles of higher animals, and in clams and oysters. It is the only form in which reserve carbohydrates are stored in animals. During starvation the liver decreases greatly in size, due to the withdrawal and consumption of the glycogen for fuel.

Cellulose constitutes the skeleton material of plants. All cell walls are made of it. In some plant tissues, as in woody stems and tree trunks, the cellulose walls are so thick and heavy that the whole tissue is hard and rigid. An especially compact and hard form of cellulose is that of nut shells and fruit stones. In the cotton plant long, fine threads of cellulose form a ball about the seeds. These threads constitute the cotton of commerce. In the flax plant, the vascular bundles of the stems when freed from the softer tissue around them, can be worked and spun into fine, hard, glossy threads. These are the linen of commerce. Paper, hemp, and jute are other cellulose products (Fig. 55).

Relation of the Carbohydrates to Each Other.—Probably some of the most important chemical characteristics of the carbohydrates are the ways in which they are converted
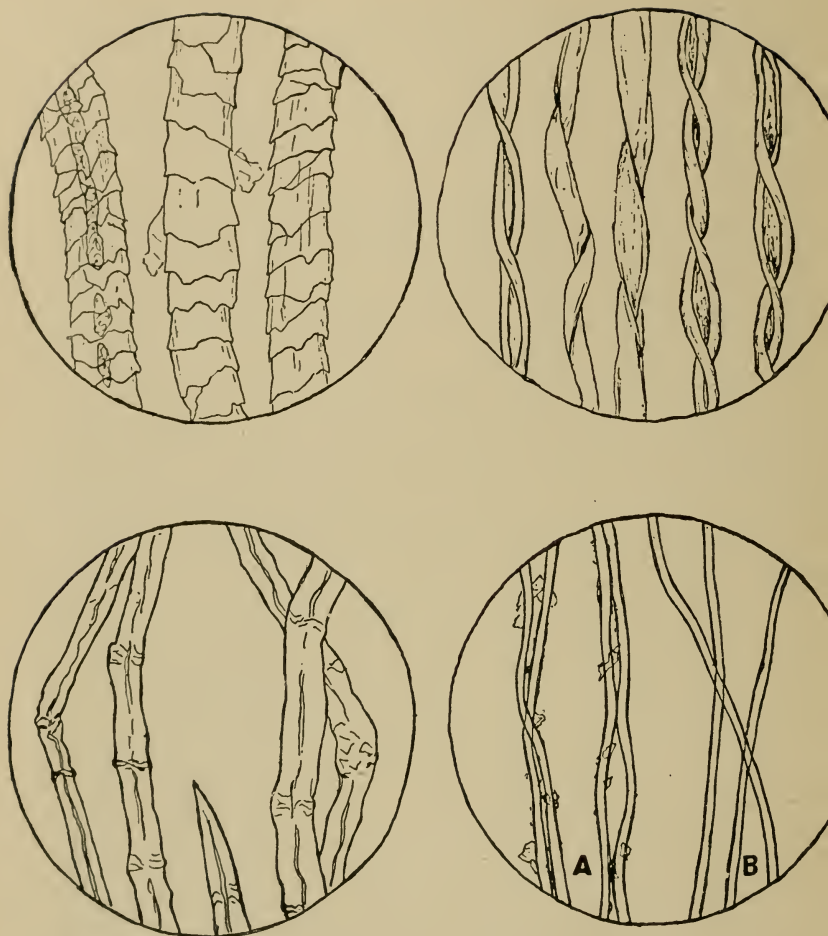
Fig. 55.—Wool, cotton, flax, and silk fibers as they appear under the microscope. All our textiles are plant products directly or indirectly, since the animals eat plants. Cotton and linen (flax) are composed of carbohydrates (cellulose) while wool and silk are composed of protein. It is often possible to tell the composition of a cloth by microscopic examination of the threads, and by chemical tests for protein and cellulose. (From Bul. 15, Minn. Station.)

into one another.   Excluding the pentoses from the discussion, we may say that the hexoses are the simplest forms of carbohydrates; they cannot be decomposed into simpler ones, and they cannot be converted into each other by simple chemical means.   However, the dihexoses, as the name indicates, are composed of two hexoses in combination.   How can a dihexose be decomposed into hexoses?   One of the simplest ways is by the action of the digestive juices.   When we eat sucrose, the pancreatic juice, by means of one of its enzymes or ferments (see Chapter XVII) converts it into levulose and dextrose, thus:

$$\underset{\text{sucrose}}{C_{12}H_{22}O_{11}} \; + \; H_2O \; = \; \underset{\text{levulose}}{C_6H_{12}O_6} \; + \; \underset{\text{dextrose}}{C_6H_{12}O_6}$$
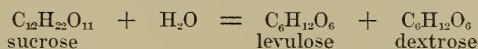
We do not know what the enzymes are chemically; all we know is what they do.   In this case the enzyme causes a molecule of water to combine with a molecule of sucrose, with the result that it is split into two molecules of hexose. Lactose is decomposed as follows:

$$\underset{\text{lactose}}{C_{12}H_{22}O_{11}} \; + \; H_2O \; = \; \underset{\text{galactose}}{C_6H_{12}O_6} \; + \; \underset{\text{dextrose}}{C_6H_{12}O_6}$$

and maltose thus:

$$\underset{\text{maltose}}{C_{12}H_{22}O_{11}} \; + \; H_2O \; = \; \underset{\text{dextrose}}{C_6H_{12}O_6} \; + \; \underset{\text{dextrose}}{C_6H_{12}O_6}$$
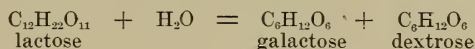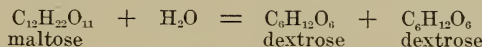
This process of combining with water is called hydrolysis. All of these disaccharides are hydrolyzed by the digestive juices before being absorbed into the blood stream.

Starch is hydrolyzed first into dextrin, then into maltose, then into dextrose.   The following equations represent the important steps in the process:

$$\underset{\text{starch}}{2C_6H_{10}O_5} \; + \; H_2O \; = \; \underset{\text{maltose}}{C_{12}O_{22}O_{11}}$$

$$\underset{\text{maltose}}{C_{12}H_{22}O_{11}} \; + \; H_2O \; = \; \underset{\text{dextrose}}{C_6H_{12}O_6} \; + \; \underset{\text{dextrose}}{C_6H_{12}O_6}$$
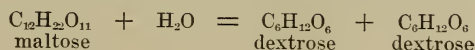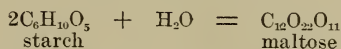
Or, these two equations can be combined into one by assuming that one molecule of $C_6H_{10}O_5$ takes up one molecule of water and becomes dextrose. Glycogen is hydrolyzed in the same way as starch. Certain forms of cellulose can be partially digested by animals but not by humans. Cellulose constitutes the bulk of the indigestible matter of foods.

Hydrolysis of carbohydrates can be brought about by agents other than enzymes. The most important of these are acids. By boiling any of the disaccharides with a little dilute acid, especially hydrochloric, they are hydrolyzed into their corresponding hexoses. Starch, glycogen, and dextrin can likewise be hydrolyzed. Corn syrup consists principally of dextrose and maltose, produced by the hydrolysis of cornstarch by means of acid. Strong sulfuric acid will attack even cellulose in such a way that it can be hydrolyzed to dextrose.

2. **Fats.**—The fats, like the carbohydrates, are compounds of carbon, hydrogen, and oxygen. But, unlike the carbohydrates, they do not contain two atoms of hydrogen for every one of oxygen. In fact, the percentage of the three elements in these two classes of compounds is very different, as is shown in the following table:

	Per cent of C	Per cent of H	Per cent of O
Carbohydrate (starch)	40	7	53
Fat (stearin)	77	12	11

Thus, there is far more carbon and hydrogen in fats; and since these are the two combustible elements, the fuel value of fats is greater than that of carbohydrates. One pound of fat is equivalent to 2.4 pounds of starch in its fuel value.

Plant and Animal Oils and Mineral Oils.—This group of compounds is often spoken of as fats and oils. What is the difference between a fat and an oil? Simply that the former is a solid, and the latter a liquid, at ordinary temperature. In our discussion here, we are excluding the so-called mineral oils. They are products of petroleum and are composed of but carbon and hydrogen, whereas fats are of plant or animal origin and are composed of carbon, hydrogen, and oxygen.

Constitution of Fats.—Each fat molecule is composed of a molecule of glycerin united to three molecules of fatty acids (so called because they occur in fats). Thus stearin, the principal fat of beef suet, can be represented as follows:

$$C_3H_5 \begin{cases} C_{18}H_{35}O_2 \\ C_{18}H_{35}O_2 \\ C_{18}H_{35}O_2 \end{cases}$$

glycerin   3 stearic acid molecules
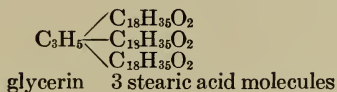
For this reason fats are called glycerides. Fats from different animals and plants differ from each other because of different fatty acids contained in them. Thus in butter-fat one of the principal glycerides is palmitin, containing palmitic acid, $C_{16}H_{32}O_2$; in olive oil it is olein, containing oleic acid, $C_{18}H_{34}O_2$; in linseed oil it is linoleic acid, $C_{18}H_{32}O_2$. Since a large number of acids are known to occur in fats, it is obvious that a great variety of fats and oils is possible.

Decomposition of Fats.—Fats are hydrolyzed by intestinal enzymes, liberating the glycerin and the fatty acids. Also, they are decomposed by boiling alkali:

$$C_3H_5(C_{18}H_{35}O_2)_3 \ + \ 3\,NaOH \ = \ C_3H_5(OH)_3 \ + \ 3\,NaC_{18}H_{35}O_2$$

stearin   sodium   glycerin   sodium stearate
                 hydroxide   (soap)
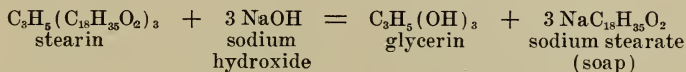
This is the process of *soap-making,* and is essentially the same for all soaps. It is seen that both glycerin and soap

are formed (Fig. 56). This constitutes almost the sole source of the glycerin of commerce, which is very extensively used in explosives and in cosmetics. Soap can be defined as the *sodium salt of a fatty acid*. The soap will differ, of course, with the kind of acid present in the fat.
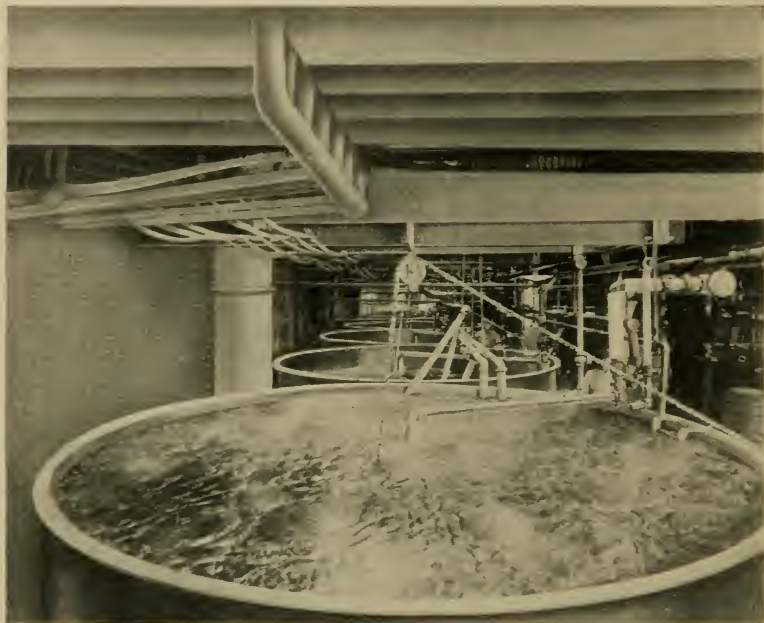
FIG. 56.—The surface of soap kettle. These kettles are several stories deep, and make 1,000 pounds of soap at a time. Fats, oils, and refuse grease of all kinds are boiled with alkali in the tub, until the fat is completely changed into soap. Salt is added to the resultant mixture, which causes the soap to rise and form a cake on the surface.

For this reason soft fats and oils will not make firm, hard soaps. When wood ashes, containing potassium carbonate, are used for soap-making, as they were by our grandmothers, the potassium salts of the fatty acids result and these are all "soft soaps."

Uses for Fats and Oils.—Probably the most important

use of these substances is as food, because of their high value in furnishing energy to the body. Another use to which enormous quantities of the cheaper fats are put is in the manufacture of soap, described above. Soap is probably the best of all cleaning agents; and it has been said that the degree of civilization of a people can be measured by the amount of soap that it uses. Linseed oil (from the seed of the flax plant), and a few others of less importance, have the property of combining with oxygen when spread out in a thin layer and of forming a hard, resistant film. It thus forms the basis for practically all paints. Some oils, such as that of the castor bean, have medicinal uses.

One of the greatest triumphs of industrial chemistry was the discovery of a way of converting vegetable oils, such as cottonseed oil, into a hard fat. The principal fatty acid in these oils is oleic, $C_{18}H_{34}O_2$; while that of the hard fat is stearic acid, $C_{18}H_{36}O_2$. In other words, they differ by two atoms of hydrogen. When a method called hydrogenation was found for causing gaseous hydrogen to unite with the oil, a solid, white fat could be made; such fats constitute practically all of our lard substitutes, and to a certain extent are used in margarines.

3. **Plant Acids.**—The sourness of fruits and other parts of plants is due to certain organic acids, compounds of carbon, hydrogen, and oxygen. The acidity of lemons, oranges, and grape fruit is due to citric acid, $C_6H_8O_7$. Malic acid, $C_4H_6O_5$, is abundant in green apples; the "sugar sand" that forms during the boiling of maple sap is calcium malate. Oxalic acid, $C_2H_2O_4$, occurs in rhubarb stems and in sorrel. Tartaric acid, $C_4H_6O_6$, is obtained from grape juice. Potassium hydrogen tartrate, $KHC_4H_4O_6$, is cream of tartar, a common component of baking powder. Lactic acid, $C_3H_6O_3$, is formed during the fermentation of silage, and during the souring of milk.

Acetic acid, $C_2H_4O_2$, is also a product of sila, fermentation; it is the acid of vinegar. Tannic acid, or tannin, is the substance used in tanning leather. It is found in many plants but is especially abundant in oak bark, oak galls, and hemlock bark. The bitter, puckering substance of strong tea is tannin.

4. **Volatile Oils.**—Under this heading will be grouped all those substances which give plants their characteristic odors and flavors. Most of them are compounds of carbon, hydrogen, and oxygen, although a few contain sulfur. The delightful odors and flavors of fruits, the delicate perfumes of flowers, the pungent aroma of pine needles, the permeating odor of garlic, mustard, and cabbage, the irritating fumes of onions, all belong to this group. Some others of great commercial importance are turpentine and camphor.

5. **Coloring Matter.**—The most important, of course, is the green chlorophyll, for this is the substance which brings about the photosynthesis of carbohydrates, as we have seen before. The coloring matter of some plants, as the indigo, the logwood, the madder, is of commercial importance as dyes.

6. **Proteins.**—These compounds are composed of the elements carbon, hydrogen, oxygen, nitrogen, and sulfur, the nitrogen being an especially characteristic element. The living protoplasm of the cells of all animal and plant tissues is composed largely of proteins. Hence, we must consider them as being of the very greatest importance in the plant and animal bodies. Muscle and nerve tissue are almost wholly protein; the white of egg is a familiar protein that is soluble in water. The casein of milk, which forms the curd in sour milk and in cheese-making, is a protein of animal origin that contains calcium and phosphorus besides the above-named five elements. The gluten of wheat, a mixture of two proteins, is one of the most

easily prepared plant proteins. By enclosing a cup of flour in a muslin bag, and carefully kneading it under water, a gummy, rubbery, elastic mass remains; this is the *gluten* to which is due the great rising capacity of wheat bread. Similar glutens can be prepared from rye, oats, and barley, but they are far less elastic than that of the wheat, which explains the fact that wheat flour is the only kind that will make a well-risen, light, porous loaf of bread. *Gelatin* is another well-known protein obtained from the bones and sinews of packing-house animals. It has a remarkable capacity for absorbing water. Even a three per cent solution of it in water is a firm jell. Gelatin is unique among the proteins in that it contains but a trace of sulfur. All seeds contain a relatively large proportion of protein, as will be brought out in the chapter on feeds.

7. Amides.—The amides and amino acids are rather simple nitrogenous compounds, found rather abundantly in leaves, seedlings, and other rapidly growing parts of plants. They are the simple building stones out of which the more complex proteins are constructed. They are important feed constituents, since a considerable portion of the nitrogen of hays is in these forms.

8. The alkaloids are compounds that are characterized by unusual and powerful activities towards the various organs of the body. Some of the better known of these compounds are the *nicotine* of tobacco; *strychnine* of strychnos wood; *cocaine* of the cocoa tree; *caffeine* of coffee and tea; *morphine* of the poppy; and *quinine* of cinchona bark. Alkaloids are not found in all plants; and when present are only in small amounts. They are nevertheless very important drugs and many plants are cultivated for their alkaloids alone.

Germination of the Seed.—There are two principal parts to a seed: (1) The embryo, or germ, which is a miniature

plant, ready to grow and develop into the roots and leaves of the new plant; and (2) the endosperm, which is a storehouse of food materials, especially starch, fat, and protein, to nourish the new plant until it can get its own sustenance from the soil and air (Fig. 57). During the

FIG. 57.—An experiment on bean seeds. The upper figures show three beans partly mutilated. The one on the left was left whole; the middle one had half of the storage organs or cotyledons removed; the other had almost the whole of the cotyledons removed. In each case the embryo, or young plant, was uninjured. The lower figures show the three plants grown from these beans. It will be noticed that although each produced a plant, the larger and more vigorous plants came from the beans having the larger amount of storage organs present. This experiment shows the purpose of the two main organs of a seed, the embryo and endosperm. The embryo is a minute, dormant plant; when given moisture and food, it develops into the roots and leaves of the new plant. The endosperm is purely a storage organ; it does not grow, but furnishes sugar, nitrogen compounds, and minerals to the growing embryo.

development of a new plant from the seed, several different stages in the process are observed: (1) The swelling of the seed by the absorption of water; (2) the development of roots; (3) the emergence of the primary leaves or

cotyledons above ground; (4) the formation of the real
or permanent leaves; (5) the shrivelling and drying up
of the remains of the seed.

Chemistry of Germination.—We are now in a position to
discuss these phenomena from the viewpoint of the chemi-
cal changes involved.   The absorption of water before any
start toward germination can be made is necessary for
two very fundamental reasons: (1) There must be an
abundance of water in the tissues for the transfer of food
materials from the seed to the growing parts of the new
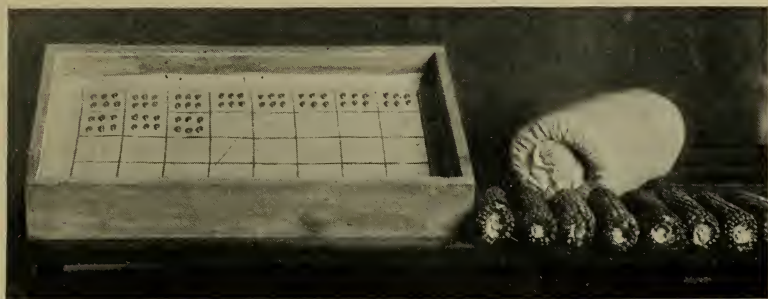
Fig. 58.—Testing seeds for germination.   Not all seeds have the power of absorbing water
and then carrying on that series of chemical changes which we call sprouting.   By wrapping
samples of each lot of seeds in a wet blanket, or by placing them in a box on wet sawdust,
the percentage of germination can be learned.

plant (Fig. 58); and (2) before the protein, starch, and
fat, which are insoluble in water, can be transferred to the
growing parts, they must undergo hydrolysis to water-
soluble compounds, and this process of hydrolysis of
course requires water.   These seed materials first absorb
water, just as a sponge does, swelling up and bursting
the seed coat.   Then by means of enzymes, or ferments,
contained in the seed, the proteins are hydrolyzed to
amides and amino acids, the starch is hydrolyzed to mal-
tose, and the fats to glycerin and soaps; and the amides,
maltose, glycerin, and soaps being soluble in water, can be
carried to the growing parts, where they are used to build

up the new tissue. These processes of hydrolysis are not completed all at once; they are performed gradually, as fast as the growing plant requires the materials.

The pushing down of new roots into the soil takes place very rapidly for a few days, because the water which the seed is able to absorb is far too scanty to supply the rapidly increasing demands of the new leaves, and more water must be obtained from the soil. Then, too, seeds are not very plentifully stocked with minerals, and a supply from the soil must be forthcoming as soon as possible.

The cotyledons, or primary leaves, are usually a part of the embryo in the seed. They are rather fleshy, and well stocked with food material, especially in peas and beans. For this reason they can quickly make their way up through the soil and into the air. Then they become green in color; and from that time on they are able to manufacture their own carbohydrates from the carbon dioxide of the air.

Very soon after this the true leaves begin to appear, and before long the new plant is able, by means of the leaves and the new root system, to obtain all its necessary food materials from the soil and the air. When this state of development is attained, the mother seed is no longer needed and it soon shrivels up and ceases to function.

QUESTIONS

1. What reasons have we for believing that plants appeared on earth before animals?
2. What products do we obtain from the earth to-day that had their origin in plant materials in by-gone ages?
3. What are our leading products that are manufactured from plant materials?
4. What is the fundamental business of agriculture? Why?
5. Make a brief outline of the ways in which chemistry applies to agriculture and to our daily life.
6. What are the raw materials out of which plants build up their tissues?
7. Compute the amount of water transpired from an acre of alfalfa, assuming 3 tons of the green crop per acre, and that 30 per cent. of this is dry matter.
8. What are the purposes of this transpiration?

9. Explain why land is sometimes left fallow in order to conserve moisture.
10. Write an equation showing the formation of starch from carbon dioxide and water.
11. Name the conditions necessary to make photosynthesis possible.
12. Make an outline of the carbohydrates, giving the formula of each, where found and their principal uses.
13. Give a definition for carbohydrates.
14. Outline the process of making granulated sugar.
15. Write equations showing the hydrolysis of starch, maltose, and sucrose, naming all substances formed.
16. Why do fats have a higher fuel value per pound than carbohydrates?
17. Explain briefly the chemistry of soap-making. What is a soap, chemically?
18. What is meant by the hydrogenation of an oil? How does it improve the oil for some purposes?
19. Name four common plant acids and state where each is found.
20. Of what chemical elements are proteins composed? Why cannot starch be used by the animal to build up muscle tissue?
21. Name three common protein substances.
22. Name four common alkaloids, and state from what plant each is obtained.
23. Write at least one page on the chemistry of a germinating seed.

LABORATORY EXPERIMENTS

65. To Study the Chemical Properties of Sugars.—(a) Heat 10 c.c. of mixed Fehling's solution * to boiling, add a few c.c. of dilute dextrose solution or diluted corn syrup (containing dextrose and maltose) and boil. The red precipitate is copper oxide, Cu_2O, and is formed by the action of the sugar on the copper compound of Fehling's solution.

(b) Repeat, using a solution of levulose, or diluted honey (containing levulose and dextrose).

(c) Repeat, using cane sugar (sucrose). If a negative test is obtained, boil 5 c.c. of a dilute sucrose solution with one drop of dilute hydrochloric acid for a few minutes, neutralize the acid with sodium hydroxide solution, and then test with Fehling's solution. What change took place in the sucrose? What is the cause of this change?

(d) Repeat, using maltose.

(e) Repeat, using lactose.

(f) What sugars give positive tests with Fehling's solution? How would you test for the sugar in a sugar beet?

66. To Study the Chemical Properties of Starch.—(a) Dissolve about 0.1 g. of starch in 20 c.c. of water by boiling. Cool, and add a drop of iodine solution † to a portion of it in another test tube. Repeat the test on the starch solution diluted with 4 volumes of water, then with 10 volumes of water.

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* Fehling's solution is made up in two parts as follows: Part A consists of 34.6 g. of copper sulfate dissolved in 500 c.c. of water. Part B consists of 173 g. of Rochelle salt (sodium potassium tartrate) and 50 g. of sodium hydroxide dissolved in 500 c.c. of water. Just before using, equal volumes of A and B are mixed.

† Dissolve 0.4 g. iodine and 1.5 g. potassium iodide in 100 c.c. water.

(b) Test the starch solution with Fehling's solution.

(c) Repeat (a) and (b) after boiling the starch solution with a drop of dilute hydrochloric acid. Explain the latter change.

67. To Show the Presence of Sugars and of Starch in Various Plant Materials.—(a) Chop up small pieces of various vegetables and fruits (or squeeze out the juices), boil them in a little water for a few moments, then test with Fehling's solution and, after cooling, with iodine solution. Make a table of the results, using + to indicate a positive test and — for a negative test.

(b) Put a drop of iodine solution on slices of raw and of cooked vegetables. Which gives a better iodine test, raw or cooked starch?

(c) Dilute some maple syrup with 10 times its bulk of water. Hydrolyze 10 c.c. of it as was done in experiment 65 (c), then test with Fehling's solution. Test another 10 c.c. portion without hydrolyzing. What it apparently the principal sugar in maple syrup?

(d) Repeat the process outlined in (c) on the extract of some slices of sugar beet. What sugar is most abundant here?

68. To Study the Properties of Fat.—(a) Smear a little cottonseed or other oil on a piece of filter paper and hold towards the light. Result?

(b) Place a few drops of oil in the bottom of each of 5 test tubes. Test the solubility of the oil in water, wood alcohol, turpentine, gasoline, and carbon tetrachloride. The latter is the safest fat solvent to use in a laboratory because it is non-inflammable.

(c) Place a few drops of oil in each of four test tubes. Add to the tubes the following: water, soap solution, 1 per cent hydrochloric acid, and 1 per cent sodium hydroxide solution. Place the thumb over the mouth of each tube and shake violently. In which cases does the oil remain in a milky suspension for some minutes? These suspensions are called emulsions. They are not true solutions of the fats as were obtained in (b). From this experiment, what two solutions would be best for washing greasy clothes and dishes? Which of these two do we use and why?

69. To Make Soap.—Place 10 or 15 g. of lard, tallow, or cottonseed or other oil in a beaker, add 100 c.c. of 10 per cent sodium hydroxide solution and boil gently. At intervals drop a few c.c. of the mixture into water. When it dissolves completely, without showing any drops of oil still unchanged, remove from the flame. The beaker now contains a mixture of soap, glycerin, water, and a little excess sodium hydroxide. Dissolve about 25 g. of sodium chloride in it, and allow to cool. The brine causes the soap to rise to the surface, where it hardens and can be removed in a cake. Examine to see how it compares with commercial soap. Dissolve it in warm water, then add dilute hydrochloric acid until the solution is distinctly acid. According to the following equations, what is the substance that rises to the top?

$$C_3H_5(C_{18}H_{35}O_2)_3 \ + \ 3\,NaOH \ = \ C_3H_5(OH)_3 \ + \ 3\,NaC_{18}H_{35}O_2$$
stearin　　　　　　　　　　　glycerin　　　　sodium
(fat)　　　　　　　　　　　　　　　　　　stearate
　　　　　　　　　　　　　　　　　　　　(soap)

$$NaC_{18}H_{35}O_2 \ + \ HCl \ = \ HC_{18}H_{35}O_2 \ + \ NaCl$$
sodium stearate　　　　　　stearic acid
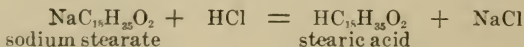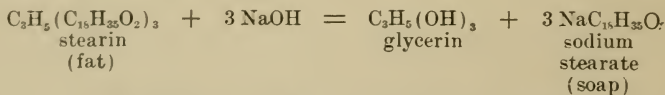
What becomes of the glycerin in the manufacture of soap?

70. To Show the Cause of the Drying of Linseed Oil.—Place a thin layer of linseed oil over the bottoms of two small bottles. Fill one bottle with CO_2 (see experiment 21) and stopper it tightly. Stopper the other bottle full of air, and let them stand for several days. Then examine them to see in which bottle the oil has dried. What constituent of the air causes the drying? How else could the drying have been prevented?

71. To Show the Reaction of Plant Juices.—With both blue and red litmus paper test the reaction of the juices of various fruits, vegetables, silage, etc.

72. To Study the Properties of Tannin.—Make a dilute solution of tannin. Add a drop of dilute iron sulfate or iron chloride solution to it. Add some of the iron solution to a strong extract of green tea. What commercial use is made of this reaction?

73. To Prepare Wheat Gluten.—Make a stiff dough out of about 25 g. of flour, let stand under water for 30 minutes, place in a muslin bag or cloth, and wash under running water until the starch is all removed. The residue is wheat gluten. Describe it. What property does wheat gluten possess which enables wheat flour to form a lighter loaf of bread than any other kind of flour? Prove the presence of nitrogen in it as was done in experiment 60. Test with concentrated nitric acid (see experiment 44). To what class of substances does gluten belong?

74. To Show the Hydrolysis of Starch Into Maltose.—Soak about 30 barley or other seeds in water for several hours, then plant in moist sand or sawdust in a box. When the sprouts are about an inch long, remove the seedlings, wash away the sand and sawdust, grind them in a mortar with about 25 c.c. of water and filter. Grind 30 unsprouted seeds in the same amount of water and filter. Test 5 c.c. portions of the two filtrates for sugar with Fehling's solution. Explain the results.

75. To Show the Presence of Fat in Seed.— (a) Place 20 g. of corn-meal in a bottle with 50 c.c. of gasoline or preferably carbon tetrachloride, and shake occasionally for an hour. Filter and evaporate some of the filtrate to dryness on a watch glass. Rub some of the residue on filter paper. Carefully heat some of it. Do these tests indicate what the substance is?

(b) Place the germ of a corn kernel in a piece of filter paper and mash it with a hammer. Explain the results.

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CHAPTER XVI

THE ANIMAL BODY

Plants the Only Feed for Animals.—Plants constitute almost the sole feed for animals, especially farm animals. If when animal tissue is eaten by other animals, we consider that this tissue was in turn produced by the eating of plants, it is readily seen that plants are ultimately the sole support of animal life. The fundamental chemistry behind this natural arrangement is simply this: The feed of animals consists principally of carbohydrates, fats, and proteins. The mineral world, the soil, the waters of the earth, do not contain such compounds, except in extremely small amounts. The carbon for these compounds must all come from the carbon dioxide of the air. Plants can feed on this source of carbon, but animals cannot. Therefore, the plants must make these compounds and then the animals must eat the plants.

We must look upon the compounds in the plant body as the raw materials from which the animal constructs its various tissues. Let us examine, then, the same series of compounds that we studied in the last chapter, just as regards their occurrence and use in the animal body.

Water.—As in plants, water is the most abundant constituent in animals. It comprises from 45 to 65 per cent of the body in most animals; in a fat hog it is only about 42 per cent. As the animal puts on fat, the percentage of water, of course, decreases. Water is not very abundant in the bones, and in the fatty tissues of hogs is as low as 7 per cent. Water plays an important part in digestion, which will be discussed later.

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Mineral Matter.—The animal body consists of from 1.5 to 5.0 per cent of mineral matter, mostly located in the bones. It is thus not as abundant a constituent as it is in the plant body. Necessarily, the elements present in the ash of animals are the same as those in the ash of plants, since the only source of these elements is the plants that are eaten. However, the animal selects certain elements and stores them in proportionately larger amounts than are found in the plants. For example, the animal requires calcium and phosphorus for the bones, and chlorine for the hydrochloric acid of the stomach; hence it must eat plant material plentifully supplied with them. On the other hand, the animal needs no silicon, which is so abundant in plants, and hence this element is practically all excreted.

Carbohydrates.—As was mentioned under this heading in the preceding chapter, the only two carbohydrates that are actually built up by the animal are the lactose of the milk and the glycogen of the liver and muscle tissue. The former, of course, furnishes feed for the young, while the latter constitutes a reserve supply of carbohydrates in the animal body. The animal forms these carbohydrates out of the various plant carbohydrates in its food.

Fats.—Animal fats represent a great variety of forms, from the thin oils of fishes to the hard tallow of beef and mutton. Their fundamental composition of glycerin and fatty acids is the same as that of the plant fats, their different characters being due to the different kinds of fatty acids. In the animals there are special kinds of fat-like substances containing phosphorus that are found especially in brain and nerve tissues. There are also protein-like phosphorus compounds that are characteristic of the liver, pancreas, and thymus gland (sweetbread). For this reason these organs are often recommended as food for children suffering with disorders of the nervous system.

The fats, in animals as well as in plants, are a reserve of nourishment in a concentrated form. They also serve as a protection against cold, as they are usually abundant just beneath the skin. The animal can convert the plant fats into its own particular kind of fat to a considerable degree. It can also convert carbohydrates into fat; hence the advice to fat people to avoid eating much sugary foods.

Plant Acids.—The acids that are so characteristic of plant juices do not have their counterparts in the animal. When these acids are eaten by animals they are usually assimilated the same as other food compounds, or are excreted in the urine in the form of salts. The only acidic fluid of the animal body is the gastric juice of the stomach; and the acidity here is due to hydrochloric acid. At times the urine may be slightly acid, due to various mineral compounds, but not to acids containing carbon.

Volatile Oils.—These do not have their counterparts in the animal world.

Coloring Matter.—The most prominent colored substance of the higher animals is the red blood corpuscles, and one of the remarkable facts in the chemistry of plant and animal life is that the green chlorophyll of plants and the colored constituent of the blood corpuscles are the same substance, except that the plant pigment contains magnesium where the animal contains iron. The functions of the two are also similar; the blood corpuscles supply the body with energy by carrying oxygen from the lungs to the various tissues, where it oxidizes the food to produce energy; and the chlorophyll in plants absorbs the energy of sunlight and enables the plant to use it in the manufacture of starch. Other animal pigments, such as those of the skin, feathers, fur, and bile are not related, so far as is known, to plant pigments; and the animal does not use the plant pigments as raw materials for constructing its own pigments.

Proteins.—The muscle tissue, the hair, skin, nails, tendons, and connective tissues, also most of the material of nerve tissues, consist of proteins. The *casein* of milk is another familiar animal protein. *Gelatin* is a protein obtained from bones. To build up these proteins the animal must eat plant proteins. In the process of digestion, as we shall see later, the proteins of the feed are broken down into the simple amino acid units of which they are composed; these are carried by the blood to the various tissues and are there rebuilt into the particular kinds of protein that are wanted.

Amides and amino acids are not found in animals except in the digestive tract and in very small amounts in the blood, where they occur only temporarily before being built up into proteins.

Alkaloids.—These are not found in any animal. If they are taken as medicine or are eaten in plants, they are eliminated through the kidneys in a short time.

General Composition of the Animal Body.—The accompanying table shows the composition of various farm

TABLE XVII

Composition of Farm Animals

	Water	Fat	Protein	Ash	Content of intestinal tract
	Per cent	*Per cent*	*Per cent*	*Per cent*	*Per cent*
Fat steer..............	45	30	15	4	6
Half-fat steer...........	51	19	17	5	8
Lean sheep.............	57	19	15	3	6
Fat sheep..............	43	36	12	3	6
Half-fat pig............	55	23	14	3	5
Fat pig................	41	42	11	2	4

animals, exclusive of the contents of the stomach and intestines. It serves as a concrete summary of the preceding paragraphs. The data are from the Rothamsted Experiment Station in England. It will be noticed that

the protein and ash content are rather constant for one kind of animal, but that the fat and water content vary in opposite directions according to the condition of the animal. The fat pig contains the least ash, protein, and water, and the most fat, of all farm animals. The carcass of the hog also produces the greatest proportion of dressed meat.

QUESTIONS

1. What is the fundamental chemical reason why animals are entirely dependent on plants for their feed?
2. Make a list of all the classes of compounds found in plants, and state specifically what the animal does with each when it is taken into the body.
3. How does plant ash differ from animal ash in composition?
4. How does chlorophyll differ from the hemoglobin of the blood corpuscles? What purpose does each serve?
5. Name 5 animal proteins.
6. Name several animal fats and give the use of each.

LABORATORY EXPERIMENTS

76. To Test for Sugar in Various Animal Tissues.— Repeat experiment 67, using various animal tissues, such as ground bone, muscle, liver, hair, etc., in testing for sugar. Results?

77. To Test for Water in Animal Tissues.—Using the apparatus used in experiment 10, test several animal tissues.

78. To Show the Presence of Fat in Lean Meat.—Using 20 g. of beef from which every visible particle of fat has been removed, test for fat according to the directions in experiment 75. Results?

THE NUTRITION OF THE ANIMAL BODY

Function of Feed.—Chemical processes are going on continually in the animal body. Some of these consist in the generation of heat to keep the body warm and to furnish the energy for doing physical and mental work. Some of them consist in the actual doing of work by muscle and nerve fibers. Still other processes consist in the transfer of material from one part of the body to another; the building up of storage tissue for fat; the conversion of feed into forms which the blood can absorb; the increase in the number of cells in the various tissues, and hence in the size of the whole body; the formation of eggs and of young and the elimination of waste products. These various activities require raw materials with which the work can be done. These raw materials constitute the feed of the animal.

In general we may say that there are four great classes of activities in the body: (1) The production of heat and energy; (2) the repair of tissue; (3) the formation of new tissue, and (4) the maintenance of health. Feed must contain the necessary raw materials to meet all four of these requirements. Each of these will be discussed separately.

(1) *The production of heat and energy* is usually brought about by the burning of carbohydrates and fats, although proteins can also serve this purpose. The amount of energy produced in an animal body is directly proportional to the amount of work performed and to the amount of heat required to keep the body warm. Thus a horse at rest does not burn up as much fuel in a day as a horse at

work. A bookkeeper requires only half as much fuel food a day as a lumber-jack or fireman. And a person living in a cold atmosphere requires more heat production in the body than a person in a warm atmosphere.

Heat is usually measured in calories. Technically the calory is the amount of heat required to raise the temperature of one gram of water one degree Centigrade. The heating value of a substance is measured by burning it in an apparatus called a calorimeter. This instrument is used for both fuels and foods. We can most easily think of calories in the following terms: An average slice of bread (1.3 ounces) furnishes about 100 calories; a glass of milk, 150 calories; a large potato, 150 calories; an egg, 80 calories; a cubic inch of butter, 160 calories; one-fourth of a pound of lean meat, 250 calories (Fig. 59).

The average man at desk work develops about 2200 calories per day; a man at moderate physical work about 2800 calories; a farmer, 3500 calories. The large amount of fat eaten by Eskimos is explainable by the great amount of heat that must be developed in the body in order to maintain its normal temperature.

(2) By *the repair of tissue* is usually meant the repair of muscle and nerve tissue, not fatty tissue; for the former are the tissues that are used in performing work, and they wear out as a consequence of this work done. Since these tissues are composed of protein substances, they require protein substances for making repairs, for the elements nitrogen and sulfur are lacking in the carbohydrates and fats. Then, too, the simple amino acids of which the proteins are composed, apparently cannot be formed by the animal body from simpler raw materials as the plants can. Therefore, the animal must take the plant proteins, decompose them in the process of digestion into the amino acids, and then recombine the latter in different proportions into the particular proteins of the animal body. For

Fig. 59.—One pint of milk—337 calories —and its energy equivalent in other food materials.

The weight for each food is given below:

	Food materials	Weight (oz.)		Food materials	Weight (oz.)
1.	Skimmed milk	32.4	13.	English walnuts	6.3
2.	Whole milk	17.2	14.	Prunes	4.6
3.	Cream	6.1	15.	Shredded wheat	3.3
4.	Cabbage	45.0	16.	Bread	4.6
5.	Onions	27.0	17.	Rolled oats	3.0
6.	Sugar	3.0	18.	Rice	3.4
7.	Honey	3.6	19.	Vanilla wafers	2.7
8.	Potatoes	17.7	20.	Bacon	1.9
9.	Grapefruit	35.0	21.	Round steak	6.2
10.	Peanuts	2.9	22.	Eggs	9.0
11.	Apples	25.3	23.	Butter	1.6
12.	Bananas	18.7	24.	Cheese	2.7

(From Farmers' Library Bul. 63, Minn. Agric. Exp. Station.)

this reason all proteins are not of equal value as food, because of their scarcity in one or more of the amino acids; for the nutrition of man, some of the best proteins are those of milk, eggs, lean meat, and beans, especially soy; the proteins of the cereal grains are of medium value; and gelatin is very poor.

Under the subject of the repair of broken-down tissue should be mentioned the repair of the bones and the replenishing of the mineral matter of the blood and of other fluids and tissues. There is a daily loss of calcium and phosphorus from the bones, of chlorine from the stomach juice, of iron, potassium, phosphorus, calcium, and magnesium from the blood. These losses must be made good by the minerals in the foods. Milk supplies an abundance of calcium, phosphorus, and potassium, but practically no iron; vegetables, especially the leafy ones, and fruit contain considerable iron, calcium, magnesium, chlorine, and potassium.

(3) *The formation of new tissues* is a process almost identical with the process of the repair of tissue just discussed. The growth of muscle and nerve tissues requires protein food; the growth of bone requires calcium and phosphorus; the increase in volume of blood requires proteins and mineral materials. Under this heading should come those special cases of the formation of new tissues and fluids, as the production of milk, the laying of eggs, and the production of young. Since each of these three products either consists of a new animal body, or is the complete feed of a young animal, it is obvious that they must contain all the chemical elements found in the animal body, and that they contain proteins built up of the proper amino acids. Hence the great necessity for providing proper feed to milch cows, laying hens, and pregnant animals.

Vitamines.—There is another factor just recently dis-

covered which enters into the question of the proper nutrition of growing animals. *Substances called vitamines, of unknown chemical composition, are necessary in food for the promotion of growth and the maintenance of good health.* There are at least three vitamines; the lack of any one of them is shown by the failure of an animal to grow, even though its food may be perfect in respect to protein, energy foods, and minerals. Milk, leafy vegetables, roots, clover, alfalfa, grass, and raw fruits are rich in vitamines; cereals and meat are poor in them.

(4) The *maintenance of good health* is a rather vague phase of the function of foods, but in some particulars it is very important. What is meant by it here is the prevention by vitamines of the so-called nutritional diseases; beri-beri, common in the Orient from the almost exclusive use of polished rice for food; scurvy, the disease prevalent among sailors and Arctic explorers, who eat practically no fresh vegetables, fruits, or milk; and possibly rickets, a common children's disease. The vitamines which prevent these ailments are abundant in milk, fruits, vegetables, and the germ of wheat and other seeds.

To summarize the above discussion of the function of food, it may be said that there are four classes of food substances: (1) Fats and carbohydrates; (2) proteins; (3) mineral elements; and (4) the vitamines.

Enzymes.—Profound chemical changes take place in food during the processes of digestion. The means by which these changes are brought about are exceedingly interesting. There are certain substances secreted by the walls and glands of the alimentary tract that have the power of inducing chemical changes to take place in the food materials. These substances are called enzymes. Their chemical nature is unknown; we do not know their formulas, nor even the chemical elements of which they are composed. They are destroyed by boiling (compare

experiment 79, page 244). The enzymes are known by
the work they do. Thus, there are starch-hydrolyzing
enzymes, protein-hydrolyzing enzymes, and so on through
a long list of substances which are acted upon by enzymes.
The saliva contains an enzyme called ''ptyalin'' which
hydrolyzes starch into maltose; the ''pepsin'' of the stom-
ach acts on proteins; the ''lipase'' of the pancreas con-
verts fats into glycerin and fatty acids.

Other living organisms secrete enzymes. Thus, when
yeast grows in a solution of sucrose, it first secretes an
enzyme which converts the sucrose into dextrose and levu-
lose; then it secretes another enzyme which converts these
two sugars into alcohol and carbon dioxide. When seeds
sprout, enzymes convert the starch into sugar; hence the
sweetness of barley malt. The ''rennet'' of the cheese-
maker is an enzyme prepared from calves' stomachs.
Several dried preparations of enzymes are now in use
in medicine.

The Digestion of Food.—The nutrition of plants and
the nutrition of animals are similar in at least one respect:
the food of each of these classes of organisms must be in
water solution. In plants it must be dissolved in the soil
water before it can enter the roots. In animals it must be
dissolved by the digestive juices before it can pass through
the walls of the intestine into the blood stream. Food has
not actually entered the tissues of an animal until it has
arrived in the blood; for the digestive tract of an animal
is simply a cavity running from end to end of the body,
and food in this cavity is not really food in the body.

Therefore, *the primary object of digestion is the
changing of food into a form which is soluble in water.*
The insoluble food is acted upon by the various fluids along
the digestive tract, each portion performing a different
part of the work. Digestion consists in chemical changes
almost entirely. The changes are brought about by the

enzymes secreted by the walls of the alimentary canal and by the liver and pancreas.

Digestion in the Mouth.—The first act in the taking of food into the body is chewing, or mastication. The object of this is to grind up the food into finer particles, and to mix them thoroughly with the saliva. The saliva contains an enzyme which converts starch into maltose:

$$2\ C_6H_{10}O_5\ +\ H_2O\ =\ C_{12}H_{22}O_{11}$$
$$\text{starch} \qquad\qquad\qquad \text{maltose}$$
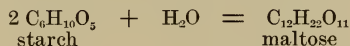
The sweet taste acquired by a piece of bread after a minute's chewing is due to the maltose formed. Proteins, fats, and carbohydrates other than starch are not affected by salivary digestion. The saliva is faintly alkaline.

Digestion in the Stomach.—The stomach has a rather strong acid reaction, due to the hydrochloric acid secreted by the stomach walls. There are two enzymes in the gastric juice; one is called rennin, or rennet, the other pepsin. The purpose of the rennin is to curdle the casein of milk, making a semi-solid out of the milk. This evidently is a provision of nature for making digestion easier. Rennet prepared from calves' and pigs' stomachs is used in cheese-making for producing the curd, and is also contained in "junket" tablets for making milk puddings. The pepsin is an enzyme which acts upon proteins. The action is one of hydrolysis (p. 213) as in the case of starch above, but the reaction is such a complicated one that it cannot be represented by an equation. The pepsin breaks down the proteins to simpler compounds, but not clear down to amino acids; that work is left for the enzymes found in the small intestine. The fats and carbohydrates are not acted upon in the stomach.

In the ruminants or cud-chewing animals, the stomach is not a simple pouch, but consists of a series of four

pouches. The food after partial chewing is first stored in the paunch. Here the saliva digestion continues, for the paunch does not secrete an acid juice, which would destroy the action of the saliva. From the paunch the animal returns the food to the mouth in the form of cuds, which are further reduced by chewing and then delivered to the other pouches.

The muscles of the stomach walls maintain a slow-churning action which serves to mix the food very thoroughly with the gastric juice.

Digestion in the Intestines.—It is here that the greater part of the digestive processes takes place. The juices of this portion of the digestive tract are again slightly alkaline in reaction. They consist of juice secreted by the walls of the intestine, of bile secreted by the liver and poured into the intestine through a duct, and of pancreatic juice, secreted by the pancreas and furnished to the small intestine by the pancreatic duct. The intestinal juice contains an enzyme for hydrolyzing whatever starch was not changed by the saliva; an enzyme for converting the maltose into dextrose; enzymes for hydrolyzing sucrose and lactose into their respective simpler sugars (p. 213); and a protein-splitting enzyme called erepsin, which continues the work of the pepsin of the stomach and decomposes the proteins into simple amino acids.

The pancreatic juice contains three enzymes: one for digesting protein, one for starch, and one for fats. The first two assist the corresponding enzymes of the intestinal juice. The fat-hydrolyzing enzyme is very important; by its action the fat molecules are split into glycerin and fatty acids. The latter combine with the sodium bicarbonate of the pancreatic juice to form soaps. Thus glycerin and soaps are the digestive products of fats. The fat enzyme is greatly assisted in its work by the bile. This contains substances which help to emulsify the fat; that is, to

break up the fat into very fine globules and to suspend these throughout the mass of digesting food. (Compare experiment 68.) In this condition the fat can be much more easily attacked by the fat enzyme.

No enzymes have been prepared from the digestive tract that will dissolve cellulose. It is very probable that the bacteria which swarm in the intestine serve to digest a portion of the cellulose into the simple sugar dextrose.

Assimilation.—From the above discussion it will be seen

Fig. 60.—A herd of beef cattle. They are chemical factories for converting the proteins, fats, and carbohydrates of the pasture grass into the corresponding compounds of their bodies, for use as human food.

that practically *the sole object of digestion is to convert the insoluble foodstuffs into water-soluble forms.* Thus the products of digestion are dextrose, levulose, galactose, amino acids, glycerin, and soaps, all of which are dissolved in the water of the intestinal contents. As fast as they are formed they are absorbed through the walls of the intestine and pass into the blood system. In this they are carried to all parts of the body, to be built up into this or that tissue or organ, or burned up for fuel, or stored for future use. The material which will not dissolve passes on down the intestinal tract and then out of the body.

In building up *muscle or nerve tissue,* the building stones are the amino acids from the digested proteins. The body selects the kinds and quantities of these building stones that are required for making a certain kind of protein tissue. Thus it is possible for us to eat egg protein, and convert it into muscle fiber; if any amino acids are left unused, they are burned up as fuel. The hen can eat the protein of grains and of meat scraps and convert them into egg protein. And a hog can eat the great variety of protein found in city garbage and convert it into pork (Figs. 60 and 61).

In building up *bony tissue,* the animal selects the calcium and phosphorus from its feed, no matter what the compounds of these two elements are in the feed, and combines them into calcium phosphate.

In storing up fat in *fatty tissue,* the animal must take the glycerin from the digested fats, together with those fatty acids now in the form of soaps, which are required to form the particular kind of fat found in that animal. Thus both a cow and a sheep may eat silage and clover hay; the sheep will convert the fat of these feeds into mutton-tallow, while the cow will convert them into butter-fat. Any fatty acids which are not used in this way are burned up for fuel. It has been found that a cow in manufacturing such large amounts of butter-fat every day must vary the selection of fatty acids to a certain degree, depending upon the kinds that are furnished in the feeds. Thus a cow eating considerable linseed meal will produce a softer fat than one eating grains which contain harder fats.

Elimination of Waste.—Just as in the burning of fuel in a stove there is waste, so also is there in the burning of fuel and in the performance of work in the body. And just as in the stove there are the ashes which are not combustible, and the gaseous products of combustion
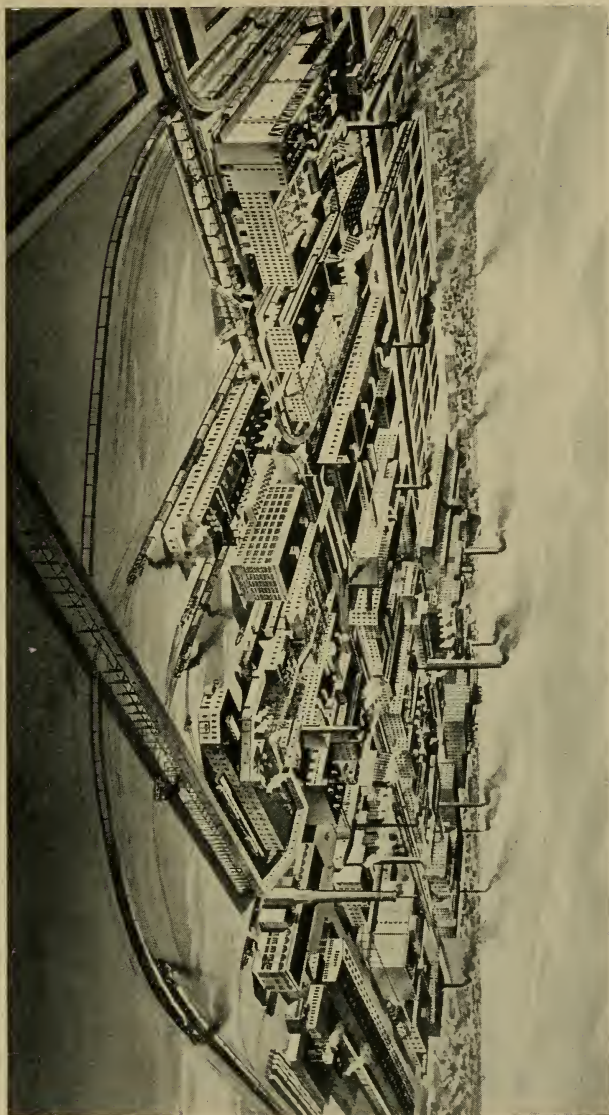
Fig. 61.—The buildings of a single company devoted to the manufacture of human foods out of animal bodies. They cover acres of space, and give employment to thousands of men. In the meat-packing industry every refuse is made to yield its by-product, thus conserving material and reducing the cost of the main products to the consumers.

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which pass up the chimney, so also are there two kinds of waste in the body. Not all of the food is digestible, hence there is a certain residue always left in the intestines which resists the action of the intestinal enzymes. This passes out of the intestines as the fæces. It contains undigested fats, carbohydrates, and proteins, as well as the products of bacterial action in the intestines—for the intestinal tract is always swarming with bacteria—and a certain amount of material from the digestive juices. The bacteria continue their action on the fæces outside the body; this is shown by the readiness with which manure rots when thrown into piles.

Products of Combustion in the Body.—That portion of the food which is absorbed into the blood and used by the tissues furnishes a variety of waste products. The fats and carbohydrates are burned by means of the oxygen taken from the lungs; *the carbon of these compounds forms carbon dioxide and the hydrogen forms water,* just as if they were burned in a fire outside the body. The carbon dioxide is thrown off through the lungs. The water is eliminated partly from the lungs, partly from the skin, and partly from the kidneys. During the performance of work by the muscle and nerve tissues and the consequent breaking down into waste products, the carbon of the proteins is partly converted into carbon dioxide and partly into compounds of nitrogen. *The nitrogen is mostly converted into urea,* CON_2H_4, which is discharged by the blood into the kidneys. This is the principal nitrogenous compound of the urine, and the one which so easily undergoes ammoniacal fermentation outside the body (see p. 186): The sulfur of the proteins appears in the urine mostly as sulfates. A variety of other compounds of nitrogen, carbon, hydrogen, and oxygen are also found in the urine, but there is no need of enumerating them here.

Summary.—From the above discussion we see that the maintenance of life and activity in the animal body involves a long series of complicated chemical changes. Only a few of these changes are known definitely; many others will no doubt be made known through future investigations. It is evident even now that the animal body is a complex chemical factory, just as the plant body is. It contains delicate adjustments for burning just the right amount of fuel necessary to maintain the temperature at a certain definite degree. It contains organs for decomposing and making soluble the raw material of the food. It contains transportation facilities for carrying the digested food material to every cell of the body. It is able to select just the right building stones for its various tissues. It can eliminate the unused portions and the waste products and thus keep the factory clear of obstructing material. The better we understand the chemistry of our bodies, and of the bodies of our farm animals, the better will we be able to care for them, and the more will we appreciate the phenomena of the living organism.

QUESTIONS

1. State the most important processes that take place in the body.
2. State what kinds of foods are used to keep these processes going.
3. What is a calory of heat? How many calories are supplied by a one-pound loaf of bread?
4. How many calories can be purchased for 10c. in the form of egg, butter, and potato, according to current prices in your neighborhood?
5. Why cannot sugar be used for building up muscle tissue?
6. What elements does bone tissue require?
7. Name the four classes of food substances.
8. What is the object of digestion?
9. Why cannot the body assimilate anything but water-soluble food?
10. What substances are formed by the digestion of protein? Of fat? Of sucrose? Of milk sugar?
11. What are enzymes? What kinds of enzymes are found in the mouth? In the stomach? In the pancreas? In the juice of the small intestine?
12. What products are formed by the combustion in the body of fats? Of starch? Of protein?
13. What are soaps?
14. What kinds of waste materials are eliminated in the fæces, in the breath, in the perspiration, and in the urine?

LABORATORY EXPERIMENTS

79. To Study the Action of Ptyalin on Starch.—Make a solution of starch by boiling 0.4 g. of it in 100 c.c. of water. Test 5 c.c. portions of it with iodine, and with Fehling's solution. Then place 5 c.c. of it in each of 4 test tubes. To tube 1 add 2 c.c. of saliva; to tube 2 add 2 c.c. of boiled saliva; to tube 3 add 2 c.c. of saliva and 1 drop of dilute hydrochloric acid; add 2 c.c. of water to the starch solution in tube 4. Allow all the tubes to stand for a half-hour in water warmed to the body temperature. Again test the contents of the tubes, using a few drops on a white dish for the iodine test and the rest for the Fehling test. What does each tube show? What happens to salivary digestion as soon as food reaches the acid contents of the stomach?

80. To Show the Action of Pepsin on Boiled White of Egg.—Boil an egg five minutes, then cut up the white, with the exception of a piece about one-fourth of an inch cube, into fine pieces. Dissolve 0.5 g. prepared pepsin scales in 50 c.c. of water. Then arrange the following test tubes:

No. 1.—Egg white + 10 c.c. of water.

No. 2.—Egg white + 8 c.c. of water + 2 c.c. of saliva.

No. 3.—Egg white + 10 c.c. of pepsin solution.

No. 4.—Egg white + 10 c.c. of pepsin solution + 2 drops of 5 per cent hydrochloric acid.

No. 5.—Egg white + 10 c.c. of pepsin solution + 2 drops of 5 per cent sodium hydroxide solution.

No. 6.—Large cube of egg white + 10 c.c. of pepsin solution + 2 of egg white drops of 5 per cent hydrochloric acid.

Keep in a place as near body temperature as possible until the next day. Examine each tube carefully and state what each tube proves. To what class of food substances does egg white belong?

81. To Prepare Gelatin From Collagen.—Collagen is the protein found in bones. Place about 50 g. of fresh ground bone in a dish. Cover with water to which a few drops of dilute hydrochloric acid have been added, cover the dish and boil gently for an hour, stirring occasionally, and replacing the water lost by evaporation. Filter, and allow to cool. Save the boiled bone for experiment 83. If the filtrate does not set, boil it down to one-half its volume and again allow to cool. To what class of food substances does gelatin belong? Test the gelatin according to the methods given in experiments 44 and 60.

82. To Show the Composition of Egg-shell.—Boil pieces of egg-shell in water for a few minutes, pour off the water, and add a few c.c. of dilute hydrochloric acid. Result? Test the resulting solution for calcium according to the directions in experiment 52. Of what compound is egg-shell mainly composed? How do chickens obtain enough calcium to form egg-shells?

83. To Show the Composition of Bones.—Dissolve the boiled pieces of bone left from experiment 81 in 50 c.c. of dilute hydrochloric acid, warming if necessary. Dilute to 100 c.c., filter, and test portions of the filtrate for calcium and for phosphorus according to the directions in experiment 52.

FOODS AND FEEDS

Coëfficient of Digestibility.—From the discussion of digestion, assimilation, and elimination in the last chapter, it is evident that the food which is eaten follows two main routes through the body. One portion, which is not digested, passes through the intestinal tract and appears as fæces; the other portion, which is digested, passes through the various tissues of the body and appears as waste products in the urine, perspiration, and breath. *The percentage of the food which does not appear in the fæces is called the coëfficient of digestibility.* For example, let us take the figures obtained by studying the digestibility of clover hay when eaten by a cow:

	Protein	Crude fiber	Starch	Fat
Amount in 100 lbs. of hay..	12.3 lbs.	34.8 lbs.	38.1 lbs.	3.3 lbs.
Amount found in fæces..	5.2	16.0	13.7	1.5
Amount digested	7.1	18.8	24.4	1.8

Coëfficient of digestibility $\frac{7.1}{12.3} = 58\%$ 54% 64% 55%.

It will be noticed that a considerable portion of each of the constituents is not digested. This will vary with the feed, and also with the kind of animal, as is shown in the table on page 246. Thus, a ruminant (cow or sheep) can digest only about 68 per cent of the protein of corn, whereas a pig can digest about 86 per cent of it. On the other hand, the cow can digest considerably more of the fat of corn than can the pig.

Such figures on digestibility are only roughly correct, since individual animals differ widely in their digestion

capacities, and the chemical methods for the determination of digestibility are not very accurate.

TABLE XVIII

Comparative Digestibility of Feeding Stuffs

	Kind of animal	Coëfficient of digestion				
		Ash	Protein	Crude fiber	Carbo-hydrates	Fat
		Per cent	*Per cent*	*Per cent*	*Per cent*	*Per cent*
Cornmeal.......	Cow....	..	68	..	95	92
	Horse...	..	76	..	96	73
	Pig.....	..	86	29	94	82
Timothy hay...	Cow....	33	47	53	62	52
	Horse...	34	21	43	47	47
Potatoes.......	Cow....	..	45	..	90	13
	Horse...	..	88	..	99	..
	Pig.....	44	85	..	98	..
Wheat bran....	Cow....	..	78	28	69	68
	Pig.....	..	75	33	65	72

The Analysis of Foods and Feeds.—It is often highly desirable to know the composition of food materials in order to compute nutritive ratios and balanced rations, and to establish the money value of a product. Therefore the chemist has learned to make certain analyses of such products, which are designed to furnish an idea of the amounts of the various classes of substances present. These analyses are usually termed as follows: (1) Moisture; (2) ash; (3) crude protein; (4) crude fiber; (5) ether extract; (6) nitrogen-free extract. The meaning of each of these terms will now be discussed.

(1) Moisture.—For this determination a weighed amount of the material in a dish is placed in an oven heated to the boiling point of water, and kept there until it no longer loses weight. The loss in weight is considered to represent the moisture content. It is evident that vola-

tile substances other than water, such as the acids in silage, and odorous substances in many plant materials, will also be lost by this drying treatment. The error due to these causes, however, is seldom more than one per cent.

(2) **Ash.**—This is determined by taking a weighed amount of the material, igniting it carefully in a special dish until there is nothing but a gray ash remaining. This is weighed, and the percentage of the whole which it represents is computed. In general, the ash represents the mineral matter of the feed; actually, however, it always contains some carbon as carbonate, which is not mineral matter; and some of the sulfur, which is a mineral element, is usually lost during the ignition. Therefore, the figures for ash give only an approximate idea of the mineral contents, but they are sufficiently exact for ordinary purposes.

(3) **Crude Protein.**—All proteins contain close to 16 per cent of nitrogen. This figure is so constant that by determining the amount of nitrogen in a sample and multiplying it by 6.25, the amount of protein is obtained. In many feeds, for example hay, a good deal of the nitrogen occurs in amids and amino acids. This nitrogen is, of course, included with that of the true protein in the analysis; hence the term, "crude protein." Since the forms of nitrogen have considerable nutritive value, there is no serious error involved in including them with the protein in our calculations.

(4) **Crude fiber** is a term applied to certain of the cellulose constituents of plant material. It is determined by boiling the material in dilute acid, then in dilute alkali, and then drying and weighing whatever portion did not dissolve in either of these. It is supposed to represent the indigestible portion of the feed. From Table XVIII, it will be seen that a considerable portion of the crude fiber is digested, however. This simply goes to show that a chemi-

cal process in the laboratory cannot imitate the processes of digestion in the animal. The only way to determine the value of a given food is to feed it to an animal and then measure the results by various means.

(5) **Ether Extract.**—Fats are soluble in ether, but are not soluble in water. Therefore, by treating a sample of material with ether, the fat is removed. Then by evaporating off the ether, and weighing the fat that is left behind, an analysis of the sample for fat is obtained. Since, however, ether extracts other substances, such as chlorophyll, wax, and certain acids, the material obtained by the above analysis is more correctly called "ether extract" than fat. In leafy material, such as hays and silage, the ether extract contains a large proportion of non-fat substances; in mill feeds and grains, the ether extract is practically all true fat. For the purposes of computing rations, however, it is sufficiently accurate to consider the ether extract as fat.

(6) **Nitrogen-free Extract.**—When the above five analyses have been made, their sum is subtracted from 100 per cent, and the difference is called the "nitrogen-free extract." It is a poorly-chosen term, as it gives no idea of the nature of the substances included under it. It is not an extract, and the crude fiber and the fat are also free of nitrogen. The water-soluble carbohydrates are the only substances not accounted for by the five preceding analyses. Instead of analyzing for starch and for sugars, which is a laborious process, these are included under the term nitrogen-free extract, assuming that in most plant materials starch and sugars do constitute the bulk of the portion remaining from the other five analyses. In feeds or foods in which sugar or starch is the characteristic constituent, such as syrups, molasses, and mill feeds, the individual carbohydrates are often determined. In this case they are not called "nitrogen-

free extracts," but are given their own designation, as sucrose, dextrose, and starch.

TABLE XIX

Composition of Some Common Foods as Percentages of the Edible Portion

	Protein	Fat	Carbo-hydrate	Fuel value per pound calories
	Per cent	*Per cent*	*Per cent*	
Apples...........................	0.4	0.5	14.2	285
Bananas.........................	1.3	0.6	22.0	447
Cherries.........................	1.0	0.8	16.7	354
Muskmelons.....................	0.6	....	9.3	180
Strawberries.....................	1.0	0.6	7.4	169
Tomatoes........................	0.9	0.4	3.9	104
Asparagus.......................	1.8	0.2	3.3	100
Beans, string, fresh.............	2.3	0.3	7.4	184
Beets............................	2.3	0.1	7.4	180
Cabbage.........................	1.6	0.3	5.6	143
Corn, green......................	2.8	1.2	19.0	455
Lettuce..........................	1.2	0.3	2.9	87
Peas, green......................	7.0	0.5	16.9	454
Potatoes, white..................	2.2	0.1	18.4	378
Potatoes, sweet..................	1.8	0.7	27.4	558
Dates, dried.....................	2.1	2.8	78.4	1575
Figs, dried......................	4.3	0.3	74.2	1437
Prunes, dried....................	2.1	0.7	73.3	1368
Beans, dried.....................	22.5	1.8	59.6	1565
Flour, wheat.....................	11.4	1.0	75.1	1610
Bread, white.....................	9.1	1.6	53.3	1200
Beef, liver.......................	20.4	4.5	1.7	584
Beef, roast.......................	22.3	28.6	.....	1576
Beef, round, lean.................	21.3	7.9	.....	694
Eggs............................	13.4	10.5	.....	672
Ham, smoked, lean...............	19.8	20.8	.....	1208
Mutton, leg......................	19.8	12.4	.....	863
Pork, chops......................	16.6	30.1	.....	1530
Oysters..........................	6.2	1.2	3.7	228
Whitefish........................	22.9	6.5	.....	680
Butter...........................	1.0	85.0	.....	3491
Chocolate........................	12.9	48.7	30.2	2768
Lard............................	....	100.0	.....	4086
Walnuts.........................	18.4	64.4	13.0	3182
Sugar, granulated................	....	.....	100.0	1815

The Composition of Human Foods.—In the accompanying table there are presented the analyses of some common foods of several different groups. It is seen that

water is a very prominent constituent in most of them, especially vegetables. In fact, the percentage of water in most vegetables is as high or higher than that in milk. Even bread contains 30 per cent or more of water.

Foods can be grouped in several ways. One way is on the basis of their origin—as vegetable foods, animal foods, manufactured foods. Another is on the basis of their characteristic constituent—as protein food, fatty food, and bulky food (containing much water and fiber). The latter basis is the one of most interest to us from the chemical viewpoint. Thus, there are the foods high in protein, such as lean meat, eggs, cheese, peas, and beans. It very often happens, as in eggs, that the percentage of protein in the whole product is low, due to the high water content; but that the greater portion of the dry matter present is protein. The commoner foods high in fat are, of course, lard, butter, olive oil, fat meats, and nuts. Some of the sugary foods are the syrups, sugar, and raisins. The starchy foods are cereal products and potatoes. The foods highest in ash are the vegetables, especially the leafy vegetables.

The Composition of Feeding-stuffs.—The following table gives the analyses of some common animal feeds. What was said concerning the classification of human foods also holds true here, although feeds are usually roughly grouped into "concentrates" and "roughages." The latter contain much water and fiber, and the former relatively little of these things. The roughages are designed to give bulk to the food, so as to increase intestinal action, to supply a certain amount of nourishment at low cost, and to furnish the vitamines necessary for the growth and well-being of the animal. The concentrates, such as grains and mill feeds, are mostly used for forced feeding, to maintain high milk flow, or to hasten the fattening process. There is less range in composition among feeds than

among the human foods—that is, there are none very high in fat, like butter and lard, and none high in protein, like cheese and lean meat.

TABLE XX

Composition of Some Common Feeding Stuffs

	Water	Ash	Crude protein	Crude fiber	Nitrogen-free extract	Ether extract
	Per cent	*Per cent*	*Per cent*	*Per cent*	*Per cent*	*Per cent*
FODDERS:						
Corn fodder (green)...	79.3	1.2	1.8	5.0	12.2	0.5
Corn fodder (field cured)	42.2	2.7	4.5	14.3	34.7	1.6
Corn silage..........	79.1	1.4	1.7	6.0	11.0	0.8
Timothy hay.........	13.2	4.4	5.9	29.0	45.0	2.5
Alfalfa hay..........	8.4	7.4	14.3	25.0	42.7	2.2
Clover hay (red)......	15.3	6.2	12.3	24.8	38.1	3.3
ROOTS:						
Turnips.............	90.5	0.8	1.1	1.2	6.2	0.2
Rutabagas...........	88.6	1.2	1.2	1.3	7.5	0.2
GRAINS:						
Corn................	10.9	1.5	10.5	2.1	69.6	5.4
Barley..............	10.9	2.4	12.4	2.7	69.8	1.8
Oats................	11.0	3.0	11.8	9.5	59.7	5.0
Wheat..............	10.5	1.8	11.9	1.8	71.9	2.1
MILL PRODUCTS:						
Cornmeal...........	15.0	1.4	9.2	1.9	68.7	3.8
Wheat bran.........	11.9	5.8	15.4	9.0	53.9	4.0
Gluten feed.........	7.8	1.1	24.0	5.3	51.2	10.6
Linseed meal (new process)..........	10.0	5.2	36.1	8.4	36.7	3.6

The legume hays are characterized by high protein, while the non-legume hays contain more carbohydrates, such as starch, pentoses, and digestible cellulose. Although the fodders and hays appear to be high in ether extract in proportion to the small amount of dry matter, it must be remembered that the ether extract of these materials is composed largely of chlorophyll and wax instead of true fats, and hence is of less value to the animal. This is shown in the table (p. 246), where the digestibility of the ether extract of the hays is far less than that of the grains.

Nutritive Ratio.—The fact was emphasized in the last chapter that the two main features in the life processes of the animal are the production of energy and the repair and formation of tissue. These two processes require different foods: the first can be supplied by carbohydrates, fats, and proteins, but the latter can be supplied only by protein. Since the protein element in foods is one of the most expensive, it is economy to supply just enough protein to maintain the growth and repair of tissue, and to utilize fats and carbohydrates for energy. Therefore, there has come into use the practice of computing the proportion of protein to non-protein constituents in a feed; this proportion is called the *nutritive ratio,* which can be defined as *the ratio of digestible protein to digestible fats plus carbohydrates.*

Computing the Nutritive Ratio.—The nutritive ratio is computed as follows: In wheat bran, for example, there is 10.2 per cent digestible protein, 39.1 per cent digestible starch, 2.1 per cent digestible crude fiber, and 2.9 per cent digestible fat. Starch is taken as the unit for energy-producing food substances; and since a pound of fat is equivalent to 2.4 pounds of starch in fuel value, the fat in the above feed must be multiplied by 2.4 in computing the nutritive ratio. A pound of digestible fiber is taken as the equivalent of a pound of digestible starch. The ratio then becomes:

$$\frac{\text{digestible protein}}{\text{digestible starch} \;+\; \text{digestible fiber} \;+\; (\text{digestible fat} \times 2.4)}$$

or

$$\frac{10.2}{39.1 + 2.1 + (2.9 \times 2.4)} \;=\; \frac{10.2}{48.1} \;=\; 1 : 4.7$$

The nutritive ratio is then 1 to 4.7. This means that for each part of digestible tissue-forming food there are 4.7 parts of digestible energy-producing food. The term "*wide* ratio" is used for one like that of oat straw, about

1 : 12, where there is a small proportion of protein; and ''narrow ratio'' for one like that of linseed cake, 1 :4, where there is a high proportion of protein.

A **balanced ration** is one in which the nutritive ratio of the feed is adjusted according to the needs of the animal. It is perfectly evident that all animals will not need the same proportion of protein to non-protein feed. Young animals, who are rapidly building up tissue, require a higher proportion of protein; that is, a narrower ration, than adult animals; and the younger the animal, the narrower the ration that is desired. Milch cows, manufacturing large quantities of protein in milk, require rather narrow rations. Horses or other animals at heavy work require narrower rations than those at light work. Balanced rations for all ages and kinds of animals, and for men at various kinds of work, have been very carefully worked out. Thus, a man at light work requires about 2500 calories of energy per day, and a nutritive ratio of about 1 to 8; a man at heavy work requires about 3500 calories, with a nutritive ratio of about 1 to 6; a young growing animal, a nutritive ratio of 1 to 5, a milch cow 1 to 6 : 5 or 7. More details concerning the feed requirements of animals can be obtained from books on stock feeding.

Since it is not the purpose of this book nor of this course to discuss the details of feeding practices, but only to give the chemistry underlying such practices and the theories governing them, we will not go more deeply into the subject of foods and nutrition, except to discuss cooking and preserving processes.

QUESTIONS

1. Why cannot a chemical analysis of the food tell us its digestibility?
2. How is the digestibility of a feed found? How do chemical analyses help in this?
3. Explain how each of the following is determined in analyzing a food: water, protein, fat, fiber, ash, and carbohydrates.

4. What is the difference between " ether extract " and " true fat " ?
5. Name several foods which have a higher water content than that of milk.
6. Why is clover hay preferable to timothy hay for milch cows?
7. What are " concentrates " ?
8. What purpose do roughages serve in animal nutrition? Name some roughages among human foods.
9. In computing nutritive ratios, why are fats and carbohydrates figured together? Why is the per cent of fat multiplied by 2.4?
10. Does a laying hen need a wide or a narrow nutritive ratio? Why?

LABORATORY EXPERIMENTS

84. To Determine the Amount of Moisture in a Food Material.— Weigh accurately about 10 g. of sliced potato, carrot, beet, cabbage leaves, apple, or other similar succulent food, in a shallow aluminum, tin, glass, or other dish that has been previously accurately weighed. Place in an oven heated to 100° C. After 6 or 8 hours, cool quickly and weigh. Place in the oven again for an hour, and reweigh. Repeat until no loss in weight occurs. The loss in weight represents moisture. Calculate the percentage of moisture in the food. Save the dried residue for the next experiment.

85. To Determine the Amount of Ash in a Food Material.— Accurately weigh a porcelain dish about 2 or 3 inches in diameter. Place in it the dried residue from experiment 84, and carefully ignite with a hot flame until only a white or gray ash remains. Cool and weigh. Calculate the amount of ash in the dried material and in the original fresh material. What elements compose this ash?

86. To Show the Presence of Fat in a Food Material.— Place about 5 g. of cornmeal in the drying oven for 2 or 3 hours. Place it in a small beaker, and cover with ether. (CAUTION: keep at least 10 feet from flames!) Allow to stand with occasional stirring for 15 minutes. Filter off the ether through a paper and evaporate it to dryness in a small beaker in a warm place away from flames. What is the residue in the beaker? How could the total ether extract in a food be determined? Make a similar extract of dried and finely chopped alfalfa hay. To what is the color due? How does the ether extract of the hay and of the corn differ?

87. To Determine the Amount of Crude Fiber in Hay.— (a) Accurately weigh 5 g. of finely chopped alfalfa hay. Place in a beaker with 200 c.c. of 1.25 per cent sulfuric acid and boil for 30 minutes. Filter through a piece of linen cloth, wash with a little water, transfer the residue to the beaker again and boil 30 minutes with 200 c.c. of 1.25 per cent sodium hydroxide. Again filter through the linen cloth, wash with hot 1.25 per cent sulfuric acid, then with hot water until the wash water is colorless. Transfer the residue to a small weighed dish, dry in the oven at 100° C. and weigh. Tne final dried residue is the crude fiber of the hay, and represents approximately the indigestible carbohydrates. It is mostly cellulose. Calculate the percentage of crude fiber in the hay.

(b) How is the protein in a substance determined? The nitrogen-free extract?

CHEMISTRY OF THE COOKING AND PRESERVING OF FOODS

So far in our discussion of the nutrients used by both man and the lower animals, no account has been taken of the effect on their composition by the ordinary processes of handling to which they are subjected. The majority of human foods are cooked; some animal feeds are cooked; and many of our foods are subjected to methods of preservation. These processes involve chemical changes; and since chemical changes in food constituents may affect their value as foods, we should have an understanding of the chemistry involved in the cooking and preserving of food. Furthermore, the clearer our knowledge of these processes is, the more effectively can we perform and control them.

Effect of Heat.—The heating of most food substances improves their palatability; that is, it imparts to them agreeable flavors and odors which stimulate the appetite and excite the flow of digestive juices. In these respects, cooking is very important and beneficial. Its effect on some of the food constituents may not be so beneficial.

(1) **The effect of heat on proteins** is to coagulate them. The familiar example of this is the boiling of an egg whereby both the white and the yolk are congealed or coagulated to a solid. All proteins are thus affected by temperatures above 150° F. There has been a great deal of discussion and experimenting concerning the relative digestibility of cooked and of uncooked proteins, some maintaining that the coagulation increases the digestibil-

ity, others that it decreases it. However, it is now pretty generally agreed that the digestibility is decreased. For this reason raw eggs are always recommended for invalids. As regards the proteins of meat, their digestibility is also decreased by cooking; but what more than offsets this is the fact that the cooking makes the meat more tender and hence can be better chewed, and that the marked improvement in flavor stimulates the digestive processes. This is often the deciding point in the question of whether hog feeds should be fed cooked or raw. Although the percentage of nutrients digested may be less in the cooked, the animals eat more, and hence it is conomy to feed the cooked food.

The coagulation of proteins by heat furnishes the explanation of certain cooking practices. Vegetables are always started cooking in boiling water, as the outer layers are then coagulated immediately, which prevents the extraction of materials from the inner tissues. Likewise, meat for soup is put into cold water and brought slowly to boiling, so as to extract as much material as possible from the meat; while meat for boiling and for roasting is seared with boiling water at the beginning.

(2) **Effect of Heat on Fat.**—If clear, separated fat, such as lard or butter, is heated rather highly, it becomes brownish in color, and gives off irritating odors. These odors are due to a decomposition product of the glycerin, called acrolein. Such a change in fats should be avoided, as it impairs their palatability. When fatty tissue is heated, the effect is to liberate the fat from the connective tissues by which it is surrounded. This is well illustrated in the "trying out" of lard; the "cracklings" are the protein connective tissues. Thus the composition of the fat in roasted meat itself is not changed; it is simply made more accessible to the digestive juices and more palatable by the changes in the tissues enclosing it.

(3) **Effect of Heat on Carbohydrates.**—Heat itself does not affect the sugars. If, however, sucrose is boiled in the presence of acids, it is hydrolyzed into dextrose and levulose (p. 213). This takes place in jelly-making. In making certain kinds of candy this reaction is also brought about by the addition of cream of tartar (a form of tartaric acid). The polysaccharides, especially starch, undergo very important changes during heating. Starch is insoluble in cold water, but on boiling it is converted into starch paste, a semi-soluble form. The raw starch is very difficult to digest by man, hence the necessity of thoroughly cooking vegetables and cereals. When starch is exposed to dry heat, as in baking bread or cake, it is converted into dextrin (p. 211), which is indicated by the brown crust. Dextrin is somewhat more easily digested than starch.

Effect of Different Methods of Cooking.—The above discussion pertains only to the individual constituents of foods. What changes occur in the composition of a food during the ordinary processes of preparation and cooking? In the first place, as regards the *boiling of vegetables,* it has been found that a large proportion of the mineral elements is removed in the peelings and that a further large proportion of minerals, besides starch, sugars, and flavoring matter, is removed during the boiling, even though the precaution is taken to sear the vegetables with boiling water at the start. The obvious ways of preventing these losses of nutrients are to boil potatoes "with the jackets on," and to boil down other vegetables to a small volume of water and then to include this water with the vegetables when served. When meat is boiled, the materials extracted are not wasted but are used in soup or in gravy.

Frying consists in heating the food rather highly in fat. If the temperature at the start is rather high, the proteins of the food, either meat or vegetable, are coagulated to

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a considerable degree, and this prevents the fat from soaking into the food. If, however, the proteins are not coagulated at the start, or if it is a cereal food comparatively low in protein, a great deal of fat is absorbed. It has been found that doughnuts which contain 5 per cent of fat before frying, contain as high as 20 per cent after frying. This absorption of fat around the particles of the food hinders mouth and stomach digestion, since the fats are digested only in the intestine, and they shield the proteins and starch from the digestive juices of these two organs.

Roasting and baking are the most ideal methods of cooking vegetables, meat, or dough foods. A dry or a moist heat can be used, a high or a low temperature, none of the nutrients are lost by extraction with water, and the flavors developed are usually superior to those in either frying or boiling.

Leavening Agents.—When a dough is made of flour, water, and salt, and then baked, the product is exceedingly hard and flinty. Examples of such foods are hardtack and macaroni. If, however, a leavening agent is added to the dough, a light, soft, porous product results. A leavening agent is one which supplies gas to the dough in the form of minute bubbles; and these bubbles, expanded by the heat of baking, puff up the dough into a porous mass. Some of the commonest materials in use for leavening purposes are *yeast, baking powder, baking soda,* and *white of egg.*

(1) *Yeast* is a mass of living, microscopic plants. It feeds on sugars, especially dextrose, and gives off as by-products of its life processes carbon dioxide and alcohol. This can be represented by the following equation:

$$\underset{\text{dextrose}}{C_6H_{12}O_6} \quad = \quad \underset{\text{alcohol}}{2\,C_2H_6O} \quad + \quad \underset{\text{carbon dioxide}}{2\,CO_2}$$

When the yeast is mixed with the dough and allowed to stand in a warm place, the yeast grows and multiplies, producing carbon dioxide gas. This gas is held by the elasticity of the gluten of the flour just as rubber balloons are expanded with air, with the result that the dough is filled with minute bubbles of gas. On baking, the bubbles

FIG. 62.—Bread-making contest at a state fair. The making of good bread is both an art and a science. The latter phase of it involves certain definite chemical reactions, which must be understood by the successful bread-maker.

expand still more, which distends the whole mass of dough to several times the original volume. The temperature of the oven vaporizes all the alcohol, so that the bread is entirely free from it. The sugar or the potato water added to bread doughs is for the purpose of feeding the yeast, as the latter feeds but very slowly on the raw starch of the dough. The baking of bread receives a great deal of attention in this country, as is evidenced by the scene in Fig. 62. (See also Fig. 63.)

Fig. 63.—The development of the flour milling industry. On the top, left, is a picture of women of Palestine grinding wheat between two stones, a method used for thousands of years. The top, right, is a Dutch windmill. Millstones are also the grinding medium here. The third picture is of one of the largest modern roller mills. The grain is reduced to flour by passing between a long series of steel rollers. By this treatment the bran layers are broken and flattened and hence can easily be sifted from the powdered starchy portions of the grain. Two or three grades of flour, and several kinds of stock feed, result from wheat by this process.

(2) *Baking powders* are mixtures of chemicals which, when they come in contact with water, liberate carbon dioxide. This serves to leaven dough in the same way as the carbon dioxide from yeast. We have seen previously (p. 132) that when an acid reacts with a carbonate, carbon dioxide is produced. This reaction is utilized in baking powders. These always contain two active ingredients: baking soda (sodium bicarbonate, $NaHCO_3$) and an acid. The acids commonly used are tartaric acid, $H_2C_4H_4O_6$; cream of tartar, $KHC_4H_4O_6$; alum, $K_2SO_4.Al_2(SO_4)_3.24H_2O$; and an acid phosphate, as NaH_2PO_4. The equation for the reaction in the case of the tartaric acid powder is as follows:

$$H_2C_4H_4O_6 \quad + \quad 2\,NaHCO_3 \quad = \quad Na_2C_4H_4O_6 \quad + \quad CO_2 \quad + \quad H_2O$$

| tartaric acid | sodium bicarbonate | sodium tartrate | | |
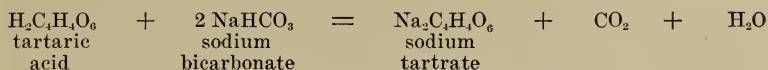
Thus there is formed sodium tartrate, carbon dioxide, and water. The sodium tartrate is a harmless salt, and is present in quantities too small to impart a taste to the dough. Similar reactions occur in each of the other types of baking powder, there being always formed carbon dioxide, water, and a salt.

It is obvious from the above reactions that definite quantities of the active substances react; thus 150 parts by weight of tartaric acid react with 164 parts of baking soda; or for each ounce of the acid there must be 1.1 ounces of soda. In this way proportions are computed for each of the types of baking powders.

In addition to the active constituents, there is always added a certain amount of "drier," usually starch, to the extent of from 10 to 40 per cent. This is to keep moisture away from the active constituents, which do not react with each other until moisture is present.

(3) *Soda,* or sodium bicarbonate, is used very often as a leavening agent in connection with foods which con-

tain acid, such as sour milk, vinegar, and molasses. The reaction is the same as in the baking powders, although the kinds of acids are different, and hence also the kinds of salts left behind in the dough.

(4) When *white of egg* is beaten to a foam, the foam consists of a small amount of protein surrounding a large amount of air in the form of minute bubbles. When this beaten egg is mixed into a cake dough and baked, the expansion of these air bubbles serves as a leavening agent. The greater the amount of egg white used, the greater will be the leavening effect.

Preservation of Food.—What is usually meant by the spoiling of food is its decomposition by the growth of microörganisms, such as molds, yeast, and bacteria. Therefore the preservation of food consists in the prevention of this action of microörganisms. In order for these organisms to thrive, several conditions must be met: (1) there must be a proper temperature; (2) there must be sufficient moisture; (3) there must be food of the proper kind and amount; (4) there must not be any poisonous chemicals present in the food. By controlling these four conditions to the disadvantage of the organisms, the latter can be prevented from spoiling food or any other product.

(1) The *temperature* at which most bacteria and molds can live and thrive is between 40° F. and 120° F. *If the temperature is above this limit, the organisms are soon killed; if it is below this limit, they are not killed but remain inactive.* It is for this reason that high and low temperatures are used to such a great extent for the preservation of food. Heating the material for some time, especially at the boiling-point of water or higher, sterilizes it, and this is the universal process in the manufacture of canned foods (Fig. 64). In the case of some organisms which form heat-resisting spores, it requires several hours of boiling to produce sterilization. Holding the
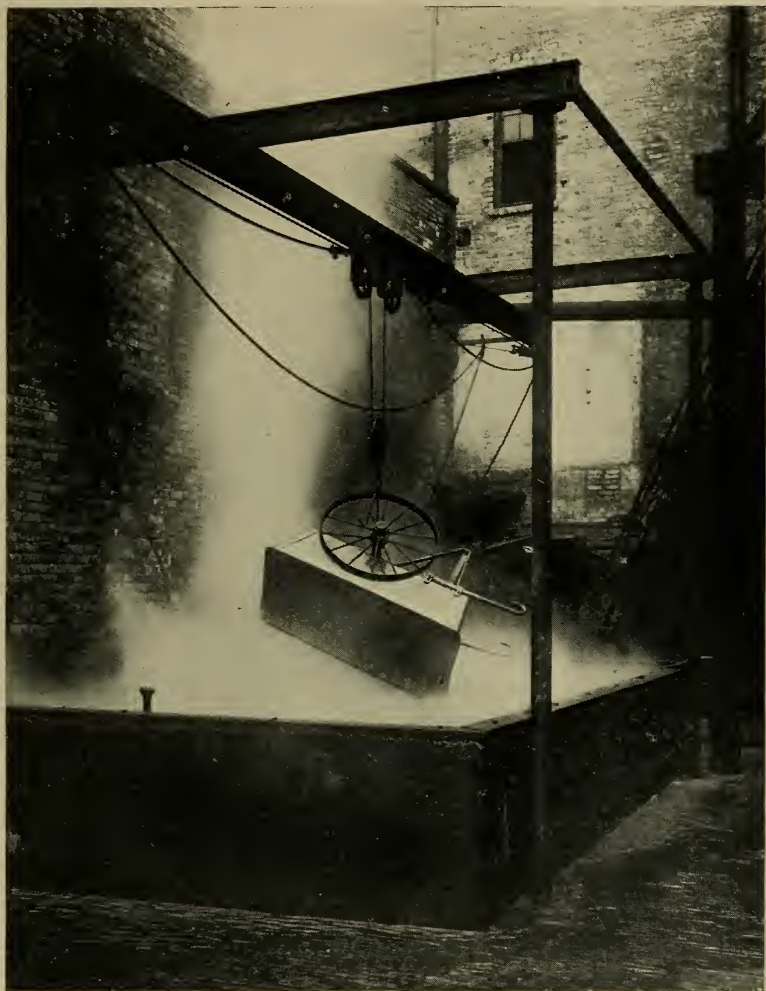
FIG. 64.—Sterilization by means of steam of trucks used in handling meat. Heat, and especially moist heat, is one of the cheapest and most effective disinfecting agents there is. Milk vessels are sometimes sterilized in a similar manner.

food in cold storage, especially below the freezing-point of water, is a very effective way of preserving foods as meat and butter, which must be kept in an uncooked, fresh

Fig. 65.—Grain elevators. These have a capacity of 112,000 bushels of grain, with machinery for aërating and stirring the grain to prevent its heating. Heating is accompanied by chemical changes which impair the germination and, in the case of wheat in particular, its milling value.

condition. Eggs, fruits, and vegetables must be kept at a temperature slightly higher than freezing.

(2) The *amount of moisture* which a food must contain before it will be likely to spoil varies greatly, according to the kind of food and the type of organisms. The

only point that interests us is the fact that when *fruits, meats, vegetables, and other food products are dried to a certain low moisture content (usually between 10 and 30 per cent), microörganisms cannot thrive in them.* Some of our great food industries are engaged in the drying of

Fig. 66.—Interior of smoke-house. After pickling the hams in salt and saltpeter, they are subjected to birch wood smoke. This smoke contains compounds somewhat similar to creosote and carbolic acid, which prevent the growth of bacteria.

prunes, peaches, apples, raisins, and other fruits; while the dehydration of vegetables has come into prominence due to a war necessity. The Indians have always preserved meat by drying. New grain must have its moisture content reduced to about 12.5 per cent, else enzyme action will develop enough heat to spoil the grain (Fig. 65).

(3) There is such a thing as a product containing *too much sugar* for the normal activities of bacteria and yeasts. Thus, when these organisms attempt to grow on jells and jams, their delicate cells are destroyed by contact with the strong sugar solutions.

(4) *Poisonous chemicals* for killing microörganisms

FIG. 67.—Curing room for hams. The fresh hams are pickled for some weeks in a solution of salt, saltpeter (sodium nitrate), and sugar. These permeate throughout the meat, and not only flavor it but preserve it. They are then smoked; the constituents of the smoke act as a further preservative.

are given various names, according to their use, as germicides, disinfectants, antiseptics, and chemical preservatives. The latter term is the usual one applied to food products. We can conveniently classify the chemical preservatives into two groups: those which are in good re-

pute, and those which are in bad repute. *Those which are in good repute* and which are widely used for the preservation of meat and fish in particular, are common salt, saltpeter, and smoke. Salt pork, corned beef, salted fish, and smoked pork and fish owe their keeping qualities to the presence of these substances, which are poisonous to bacteria (Figs. 66 and 67). The smoke contains chemical substances related to creosote. Other harmless preserva-

Fig. 68.—A pair of modern silos. In these huge containers the green plant undergoes fermentative changes which preserve the fodder in a fresh and succulent condition for many months.

tives are vinegar for pickles and water glass for eggs. Silage is green fodder which is kept from spoiling by the development of acids and the exhaustion of the oxygen in the silo (Fig. 68).

Certain other chemicals, which are not only poisonous to the bacteria, but also to man if taken in anything but small quantities, are sometimes used to prevent the spoiling of food. The commonest preservatives of this class are formaldehyde, boric acid, benzoic acid, and sulfites. Practically all state governments as well as the national

government have stringent laws against the use of such preservatives in food products. The objection to them is not based so much on their harmfulness to those who consume them, but on the fact that their use enables the manufacturer to use spoiled and inferior raw materials, and to be careless in his methods of handling them. In order to protect the purchaser, the food laws require that if a preservative be used in a food, it shall be so stated on the outside of the package. One of the most frequent violations of this law is the use of formaldehyde in milk to prevent its souring.

QUESTIONS

1. Why is it desirable to eat foods which have agreeable flavors and odors?
2. What does heat do to proteins? How does this affect their digestibility?
3. In broiling a steak, why should it be seared on both sides on a hot griddle at the start?
4. Write an equation showing the effect of boiling sucrose in the presence of an acid.
5. Why should all vegetables be thoroughly cooked?
6. What is the brown substance which forms on toast?
7. Why should potatoes be boiled " with the jackets on "?
8. What is yeast? Name two great industries which make use of yeast fermentation (see equation pp. 74 and 258).
9. What three constituents do all baking powders contain? What is the purpose of each?
10. Devise a recipe for mixing up five pounds of tartaric acid baking powder, allowing for 15 per cent of starch in it.
11. What takes place when sour milk and soda are used in the same batter? What product is left behind?
12. Name the four factors which govern the life activities of bacteria and molds. Describe how apples might be preserved by five different methods.
13. State two objections to the use of chemical preservatives in the canning of food.

LABORATORY EXPERIMENTS

88. To Show the Effect of Heat on Protein.— Dissolve the white of an egg in 5 times its volume of water. Divide into two portions. Heat one portion in a test tube or small beaker very slowly, shaking frequently. When a cloudiness in the solution shows that the protein is being coagulated, remove from the flame and quickly take the temperature of the solution. Now heat the other half of the egg-white solution to a temperature of 5° C. less than the temperature at which the first portion coagulated. Keep it at this temperature for some time. Why does not the protein coagulate? What do you infer from this concerning the condition of the pro-

tein in the center of a piece of " rare " roast beef? Which are probably the more digestible, coagulated or uncoagulated proteins?

89. To Show the Effect of Dry Heat on Starch.—Carefully remove about 0.5 gram of the very outer crust of some very well browned bread. Crumble it up fine and suspend in 100 c.c. of water. Suspend a similar amount of the crumbled white portion of the bread in 100 c.c. of water. Boil both solutions a few moments, cool and add 3 drops of iodine solution to each. Judged by the intensity of color in the two solutions, has the browning of the bread destroyed any of the starch? Into what was it converted?

What is the effect of moist heat (that is, boiling) on raw starch? Compare with experiments 66 (a) and 67 (b).

90. To Prepare a Baking Powder.—Dissolve a very small amount of tartaric acid in a few c.c. of water. Does any action take place? Dissolve a similar amount of baking soda in water. Any reaction? Thoroughly mix equal portions of dry tartaric acid and baking soda, and then add the mixture to water. Reaction? Add the solution of tartaric acid prepared in the first place to the solution of baking soda. Reaction? What gas is given off? Test it with a burning match. Test it with a drop of clear limewater on the end of a clean glass rod. What do these tests show? What is a baking powder? Why must moisture be kept away from such powders? How are commercial baking powders usually kept dry? Compare the chemistry of baking powders with that of chemical fire extinguishers.

91. To Show the Effect of Temperature and of Preservation on the Keeping Qualities of Milk.—Arrange the following series of test tubes:

(a) 10 c.c. of milk.
(b) 10 c.c. of milk stood in a beaker of boiling water for 10 minutes.
(c) 10 c.c. of milk and one drop of hydrogen peroxide.
(d) 10 c.c. of milk and one drop of dilute formaldehyde solution.

Place these tubes in a warm place (25 to 30° C.) and examine at frequent intervals to see when they become curdled. Label a fifth tube (e) and put in it 10 c.c. of fresh milk, store in a cool or cold place, and note the progress of souring. Name three effective ways of preserving milk shown by this experiment. Which of these is the least desirable, and why?

MILK AND ITS PRODUCTS

MILK is the secretion of the mammary glands of mammals. It is elaborated as a food especially adapted to the needs of young animals. A brief discussion of its chemical composition and characteristics will serve to impress us with the fact that it is a most remarkable fluid, and will also show why the civilized nations have developed the dairy cow and her products so extensively (Fig. 69).

Secretion of Milk.—Milk is a mixture of many different substances. These substances are manufactured by the cells of the udder, out of raw materials supplied by the blood. Every judge of dairy cows knows how richly the udder is supplied with blood-vessels. This is very necessary, for the milk is not manufactured by the udder continuously between milkings, but is formed mostly during the operation of milking; hence the blood must bring to the cells an abundant supply of raw materials. Now the interesting and remarkable fact about this formation of milk from the substances in the blood is, that the milk constituents themselves are not found in the blood, nor in fact in any other part of the body. Thus, the fat of milk is different entirely from the fat of any other part of the body; the casein is different from the blood proteins; and milk sugar is not found in any other part of the body or in any plant body. Therefore the cells of the mammary gland must take the fatty acids and glycerin which the blood brings from the digested food and remake them into the particular kind of fat wanted in the milk. And they must sort out the amino acids of the blood and com-

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Fig. 69.—A herd of dairy cattle. The dairy cow is the most efficient of all farm animals in converting rough feed into human food. She eats large quantities of grass, corn stover, and mill feeds; she converts the protein of these substances into the casein of her milk, out of which we make cheese; the fats she makes into butter-fat; and the carbohydrates into milk sugar, which in turn forms the flavoring matter of butter during the ripening of cream.

bine the proper ones into casein; and so on for all the constituents of milk. Furthermore, the mineral elements in milk are in entirely different proportions from what they are in blood.

Constituents of Milk.—The principal constituents of milk are *water, sugar, proteins, fat, and ash.*

The *sugar* found in milk is *lactose.* This belongs to the dihexose group, having the formula $C_{12}H_{22}O_{11}$. When it is acted upon by the digestive juices it is hydrolyzed as follows (see p. 213):

$$\underset{\text{lactose}}{C_{12}H_{22}O_{11}} \quad + \quad H_2O \quad = \quad \underset{\text{dextrose}}{C_6H_{12}O_6} \quad + \quad \underset{\text{galactose}}{C_6H_{12}O_6}$$
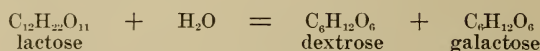
Milk sugar is very much less sweet than cane sugar. It is not fermented into alcohol and carbon dioxide by ordinary yeast, although it is by certain other yeasts. It used to be thought that lactose is more readily digested and less easily undergoes acid fermentation in the digestive tract than cane sugar. This does not appear to be the case, however.

Lactose is very quickly converted into lactic acid, $C_3H_6O_3$, by the so-called sour milk bacteria. This acid causes the curdling of milk. In fresh milk there is about 0.2 per cent lactic acid; when the acidity has reached 0.4 per cent, a sour taste is perceptible; at 0.7 per cent curdling can be seen. The acidity may go as high as 2.0 per cent after long souring.

The *proteins* of milk are *casein* and *albumin.* The casein gives milk its chalky-white appearance. Besides being composed of carbon, hydrogen, oxygen, nitrogen, and sulfur, like all other proteins, casein is unusual in that it is also combined with calcium and phosphorus.

Casein can be made to precipitate out, or curdle, by two different methods. One of them is by adding an acid to the milk. This takes calcium away from the casein, and the

latter then becomes insoluble. This is what happens during the ordinary souring of milk. The other way of curdling milk is by means of the enzyme called rennin, found in the stomachs of all animals. In this case calcium is not abstracted from the casein; in fact, calcium must be present in the milk or curdling will not take place. Why nature has provided for the curdling of milk as soon as it enters the stomach is not known. It no doubt has something to do with the ease and completeness of digestion. Rennin is prepared commercially from the stomach of calves, pigs, and sheep, for use in cheese factories.

The *albumin* in milk is in solution and hence is invisible. It is not coagulated by acids nor by renning, but it is coagulated by heat. Thus it always remains in the whey, where it can be seen if the whey is boiled. The albumin does not contain calcium or phosphorus.

Butter-fat is a mixture of a number of fats. Some of these fats are found in no other plant or animal fat, and give to butter-fat some of its characteristic properties, such as its power to absorb odors, and its penetrating odor when rancid. These particular fats are lacking in oleomargarine, and this constitutes the principal means of distinguishing butter from margarines.

Fat occurs in milk in minute globules. These globules are larger in the milk of Jerseys and of Guernseys than in that of other breeds, hence the cream in the milk of these breeds rises more quickly than in the latter.

The *ash* of milk consists of the following eight elements: calcium, magnesium, potassium, sodium, sulfur, phosphorus, chlorine, and a trace of iron. Whole milk contains about 0.72 per cent of ash. Of this, about one-fourth is potassium, one-fourth phosphorus, one-fifth calcium, and smaller amounts of the other elements. These are the mineral elements necessary for animal life, as we saw in Chapter XI; we would therefore expect milk, which

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serves as the sole food for young animals, to contain them.

Composition of Milk.—The amounts of the above constituents in cows' milk vary considerably, according to breed, period of lactation, age, and individuality. The following is the average composition, compiled from thousands of analyses of American milks:

	Per cent		Per cent
Water	87.1	Casein	2.6
Sugar	5.1	Albumin	0.6
Fat	3.9	Ash	0.7

The fat is the constituent most subject to variation, it being frequently as low as 2.0 per cent and as high as 9.0 per cent.

Effect of Breed on Composition of Milk.—One of the factors affecting the composition of milk is the breed of animal. The Jerseys and Guernseys produce milk that averages over 5.0 per cent of fat, and 3.0 per cent of casein; while the milk of Holsteins usually contains not over 3.4 per cent fat and 2.2 per cent casein. Since the latter breed produces greater quantities of milk per cow, however, the total output of milk solids does not vary a great deal among high-grade animals of the various dairy breeds.

Individual Variation Among Cows.—What is of more significance to the dairyman than breed is the production by the individual cow. No two cows of the same breed and age, and eating the same food, will produce milk of the same composition. And the differences between them may be very great, the total solids being as low as 9 per cent and as high as 15 per cent. Usually feed and care will help the poor producer to only a limited extent. Hence the practice that has come into vogue of testing the composition of the milk from individual cows, especially for the fat content, so as to provide a basis for weeding out the unprofitable ones (Fig. 70).

Effect of Stage of Lactation.—The period of lactation affects the composition of milk to a considerable degree. The first milk, or colostrum, is very high in albumin, containing 12.0 per cent or more of this protein. The colostrum gradually changes to normal milk in eight or ten days. During the first month the fat and casein are somewhat high; these show a decrease during the second and

Fig. 70.—Outfit for testing milk on the farm.   With this equipment and a bulletin of directions from the state agricultural experiment station, a farmer can test the milk of each cow for butter-fat at frequent intervals, and thus determine which animals are profitable and which are not.   (From Circular 78, Purdue Sta.)

third months, followed by a gradual increase up to the end of the lactation period.

Influence of Feed on Composition of Milk.—The feed may affect the quantity of milk produced, but it exercises but very little effect on the composition. This has been proved again and again at various agricultural experiment stations, since the proposition of feeding cows in a particular way in order to increase the per cent of fat is an ever-recurring one. In one experiment a herd of cows were poorly fed for a while, and then liberally fed. During the latter period the total flow of milk was greatly increased, but the percentage of constituents remained practically the same, except for a very slight increase in fat. As has been mentioned before, the fats in certain kinds

of feed produce butter-fats of different degrees of hardness. Thus cottonseed meal produces a harder butter-fat, of lower melting point, while gluten feed, rich in oil, produces a softer butter. Similarly, early spring grass produces softer butter.

Butter.—The first step in the process of making butter from milk is the separation of the cream. This is now usually done by centrifugal force in "separators." The amount of fat in the cream can be varied at will, but for butter-making it is usually between 20 and 35 per cent. Only 0.05 to 0.10 per cent of fat is left in the skim-milk.

The fresh cream is "ripened" or allowed to sour, in order to produce the characteristic butter flavor and odors. This ripening is brought about by the ordinary lactic acid bacteria found in milk. In many creameries the cream is first pasteurized by heating to 145° F. for 15 or 20 minutes to kill most of the harmful bacteria. It is then inoculated with a pure culture of sour milk bacteria called a "starter" and allowed to ripen. When the acidity has risen to about 0.5 per cent the cream is churned. This brings about a separation of the fat from the rest of the milk. When the fat grains have reached a certain size the buttermilk is drained away, the butter washed, and then salted. This process leaves but a small amount of milk solids (mostly casein and lactose) in the butter. The latter is composed on the average of about 80 to 85 per cent of fat, 12 to 15 per cent of water, 0.6 to 1.4 per cent of casein, and 2.0 to 4.0 per cent of salt.

Cheese.—This dairy product has been made since ancient times from the milk of cows, goats, sheep, camels, and mares. The essential processes are the curdling of the milk by means of the enzyme rennin, which is extracted from calves' stomachs, the separation of the curd from the whey, the pressing of the curd into molds, and then ripening the product for some weeks or months. It is

during this ripening that the characteristic flavors are developed, due to the action of various bacteria and molds. The principal variation in the composition of cheeses is in the fats, as they may be made from milks of widely differing fat content, from partially skimmed milk to cream. As the casein coagulates, the curd brings down with it most of the fat, together with a little of the milk sugar and ash.

Comparison of Dairy Products.—For the sake of comparing the various products made from milk, the accompanying table has been prepared. The fact of greatest significance seen in these data is that all of these products are very high in food substances. And this, together with the fact that the fats, proteins, and carbohydrates of these products have very high coëfficients of digestibility, shows that the reputation of these substances as foods of the highest order is thoroughly justified.

TABLE XXI
Composition of Various Dairy Products, Stated in Percentage

	Water	Fat	Protein	Sugar	Acid	Ash	Salt
Whole milk	87.0	3.9	3.2	5.0	0.2	0.7	. . .
Skim-milk	90.1	0.1	3.7	5.1	0.2	0.8	. . .
Cream	76–54	15–40	2–3	3–4	0.1–0.5	0.5	. . .
Butter	15.0	81.0	1.1	trace	trace	0.05	2.0
Buttermilk	90.3	0.5	3.4	4.4	0.7	0.7	. . .
Cheese	35.0	33.0	28.0	0.5	0.3	. . .	3.2
Cottage cheese . .	80.0	0.1	18.0	1.0	0.5	. . .	0.4
Condensed milk (sweetened) . . .	25.7	10.7	8.5	53.8	. . .	1.3	. . .
Milk powder	4.0	3.0	32.3	53.8	. . .	6.9	. . .

QUESTIONS

1. Describe the physiological processes involved in making milk out of silage, discussing each of the constituents of the milk.
2. What is lactic acid? How is it formed in milk?
3. Describe the difference between casein and milk albumin as regards their chemical and physical properties.
4. How does butter-fat differ from other animal fats?

5. What chemical elements are found in milk? Why do these elements and not others occur in milk?

6. State the factors which may affect the composition of milk, and tell in what way they do so.

7. Of what value are the analyses of the milk of individual cows?

8. What chemical changes take place during the ripening of cream?

9. What is pasteurization? What is its purpose?

10. What weight of food materials is there in a gallon of buttermilk?

11. What is the effect of rennin on milk? Of acid?

LABORATORY EXPERIMENTS

92. To Separate Pure Fat From Milk.— Moisten about 30 g. of dry, clean sand with about 10 g. of whole milk or cream, and place in the drying oven at a temperature not above 100° C. When dry, crumble the sand, place in a filter paper in a funnel, and drip through the sand about 50 c.c. of gasoline, ether, or carbon tetrachloride. Catch the filtrate in a beaker and evaporate off the gasoline. What is the residue? How could this method be used to determine the amount of fat in milk? What method is usually used? How is fat usually separated from milk?

93. To Separate the Casein From Milk.— (a) Dilute 10 c.c. of skim-milk with 20 c.c. of water, add 1 c.c. of 10 per cent acetic acid, stir and warm gently until a curd begins to form. When it is well flocculated, filter through filter paper. Wash the curd with a little water. Save the filtrate, which should be very clear, for the preparation of milk albumin in the next experiment. Examine the curd. Partially dry some of it, then grind with some sodium hydroxide, place in a test tube, and heat. The smell of ammonia indicates that the curd, or casein, is what kind of a compound?

(b) Dilute 10 c.c. of skim-milk with 20 c.c. of water, warm to about 30 degrees, and add a little rennin solution or a portion of a " junket " tablet. Let stand until the curd is well formed. Filter, and test the curd in the same way as that in (a). Save the filtrate for the next experiment.

(c) What kinds of cheese are made by method (a) and by method (b)?

94. To Prepare Milk Albumin.—Milk albumin is the second most important protein of milk. It is not coagulated by acid or by rennin, but is coagulated by heat the same as egg albumen in experiment 88. Heat the filtrate from the casein in (a) and (b) in experiment 93, to boiling, and boil slowly until the curd forms. This curd is the coagulated milk albumin. It can be collected on a filter paper and tested for protein, the same as the casein. Save the filtrate for the next two experiments. In the making of cheese is this protein wasted?

95. To Show the Presence of Sugar in Milk.—Heat a portion of the filtrate from the milk albumin to boiling, add a few c.c. of mixed Fehling's solution and boil. What sugar is present in milk? What is it used for commercially? What do you think would be a cheap source of it?

96. To Show the Presence of Mineral Elements in Milk.—Evaporate the rest of the filtrate from the milk albumin to a small volume, transfer to a porcelain dish, evaporate to dryness, and ignite to a white ash. What elements make up this ash? Test for calcium and chlorine in this ash by the methods used in experiments 49 (d) and 52 (c).

97. To Separate the Constituents of Butter.—(a) Fill a large test tube about two-thirds full of butter. Set it in a beaker of water and heat

the latter to about 50° C. until the butter is completely melted. Allow to cool slowly. Several different layers of substances will be observed. The fat is on top, and if the cooling has been sufficiently slow, it will be seen to consist of both solid and liquid portions. The liquid is mostly olein, and the solid mostly palmitin. Below the fat is a watery layer, containing casein in suspension and salt in solution. How did the two latter substances get into the butter?

98. To Show the Heat Test for Butter and for Oleomargarine.—Heat small samples of oleomargarine and of butter in spoons over the flame. Note the difference in the way they boil and in the amount of foaming.

99. To Show a Difference in the Kinds of Fat in Butter and in Oleomargarine.—Melt a piece of butter the size of a hickory-nut in one test tube, and a similar quantity of nut-margarine in another. Allow the water and curd to settle a moment, then pour the clear fat into beakers, leaving the curd behind. Add 30 c.c. of 10 per cent sodium hydroxide to each and boil gently for 30 minutes, replacing water lost by evaporation. Cool somewhat, then add 20 per cent sulfuric acid until the solution is acid. Carefully note the odor of the fatty acids which rise to the surface in each case. Of what does that of the butter remind you? The acids which can be smelled are volatile. Butter-fat contains considerable volatile acids, while oleomargarine fat contains very little. Write equations showing the reactions that took place in this experiment. (Compare experiment 69.)

CHEMISTRY OF CLEANING

The higher the stage of civilization of a people, the more particular are they concerning the removal of dirt of various sorts from their bodies, their clothes, and their dwellings. And the greater the care exercised in maintaining cleanliness, the greater is their consumption of soap. For soap is rightly called the universal cleansing agent. Since soap has been known for a very long time, and since it is now a cheap and common article of daily use, it is almost inconceivable that only a hundred years ago in Europe bathing was a luxury, and enjoyed only by the nobles and the wealthy. During the World War, soap became one of the scarcest of commodities in Europe, since all possible fats had to be converted into human foods. In the United States we are blessed with an abundance of cheap sources of soap-fat, which makes it possible for soap to be enjoyed by the richest and the poorest.

Principles of Cleansing Processes.—There are in general three distinct methods of cleaning:

1. The removal of dirt by emulsifying it.
2. The removal of the dirt by dissolving it.
3. The destruction of the color by chemical means.

1. *An emulsion is a suspension of drops of one liquid in another in which it is not soluble.* For example, kerosene is not soluble in water; but if it is shaken violently with water, minute drops of it remain suspended, giving the water a milky appearance. This is an emulsion of kerosene in water. A fat or oil can be similarly emulsified in water. Such emulsions are not permanent, however; the kerosene or the oil gradually rises to the surface, the

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fine drops unite into larger ones, and soon there is no longer an emulsion. *If soap is dissolved in the water, the emulsion will remain a great deal longer.* And it is this remarkable power of soap to make emulsions and to maintain them which makes it such a valuable cleaning agent. For most forms of dirt in clothing and on the skin do not dissolve in water; but when soap is added to the water, the material of the dirt is emulsified, and thus easily washed away in the soap solution. Then, too, soaps are weak alkalies; and alkalies exert a solvent action on grease.

2. The *dissolving* of stains and dirt depends upon the fact that some liquid can be found in which the stain is soluble. Thus some fruit stains will dissolve in water; grass stains will dissolve in alcohol; grease stains will dissolve in gasoline, chloroform, turpentine, and other similar solvents; and paint stains will dissolve in turpentine. If the nature of the spot is known, the correct solvent can be applied at once. If it is not known, it is always best to try water first; cold water, in plentiful amounts, then warm water. If, after drying, the stain is still visible, gasoline, carbon tetrachloride (carbona), or benzine can be used. The stain should be rubbed with a cloth from the outside towards the center so as to avoid spreading the stain.

3. If stains cannot be removed by emulsification or by dissolving out with a solvent, they can often be *bleached to colorless compounds by chemical means.* Thus oxalic acid removes many ink stains by forming colorless compounds; bleaching powder removes fruit stains; straw hats are bleached by sulfur dioxide (see p. 135); yellowness in cloth is readily removed by the action of sunlight; hair and other substances can be bleached by means of hydrogen peroxide. Usually the nature of the stain should be known before a bleaching agent is tried. Also, the possible effect of the reagent on the material to be bleached should be

known, or found out by trial on a small sample. In the case of colored fabrics this is important, for an acid may destroy the color, and an alkali restore it, or *vice versa*. Ammonia water and weak vinegar can be used for alkali and acid, respectively, in such cases.

Soaps.—As has been discussed in a former chapter, soap is made by boiling a fat with an alkali:

$$C_3H_5(C_{18}H_{35}O_2)_3 \quad + \quad 3\,NaOH \quad = \quad C_3H_5(OH)_3 \quad + \quad 3\,NaC_{18}H_{35}O_2$$

<div align="center">
stearin   sodium   glycerin   sodium
hydroxide   stearate
(soap)
</div>
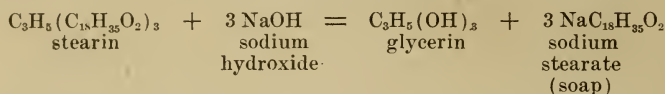
Thus one molecule of fat (in this case stearin of beef or mutton suet) combines with three molecules of alkali to produce one molecule of glycerin and three molecules of soap. *Soap is the sodium or potassium salt of a fatty acid.* The commonest fatty acids in fats that can be used for soap-making are palmitic, stearic, and oleic; as the latter makes a soft soap, only a small amount of it can be present in a soap fat. *Potassium soaps are soft.* The soft soap of our grandmothers was made from potassium carbonate leached from wood ashes.

The above equation tells us that definite amounts of fat and alkali react. If more than the correct amount of fat is used, it will remain as fat in the soap and make the latter greasy. If more than the correct amount of alkali is used, the resultant soap will be alkaline, and it will have a caustic action on skin and on cloth. Therefore, the process of soap-making is subjected to careful chemical control.

Glycerin is always formed as a by-product. It is removed from the soap, as it is a valuable raw material for explosives and other commodities. In the so-called glycerin soaps a portion of the glycerin is put back into the soap.

Composition of Commercial Soaps.—Although soap is essentially a mixture of sodium salts of fatty acids, there

are numerous other substances found in the soaps on
the market. *Water* constitutes from 15 to 35 per cent of
them; the softness of many soaps, and the rapidity with
which they wear away in water, are due to their high water
content. Practically all soaps contain *perfume*. In some
cases this is added to masque the odor of rancid fats used
in the manufacture, but more often the perfume adds to
the attractiveness of the product, and certain perfumes
act as preservatives for the soap by preventing moldiness.
The most common adulterants of soaps are the so-called
fillers. These are usually heavy, inactive substances such
as chalk, infusional earth, sodium sulfate, and water-
glass, added to the soap to give it weight. *Alkaline* soaps
are those which contain sodium or potassium carbonate.
These are undesirable constituents, since they are harsh
on the hands and on fabrics. Very often substances of
supposed medicinal value are added, such as tar and car-
bolic acid. Again, there are many *scouring soaps and
scouring powders* on the market; these consist largely
of sodium carbonate and soap, with pumice as an abrasive.
Sometimes borax, naphtha, or kerosene is added to soap;
these increase its cleaning power.

Hard Water.—Hard water is one of the most frequent
causes of the uneconomical use of soap. The calcium and
magnesium of the water combine with the soap to form
lime soap and magnesium soap :

$$2\,NaC_{18}H_{35}O_2 \; + \; MgSO_4 \; = \; Mg(C_{18}H_{35}O_2)_2 \; + Na_2SO_4$$

| soap | magnesium sulfate | magnesium stearate | sodium sulfate |
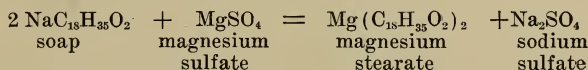
The magnesium and calcium soaps are insoluble, and form
the sticky, gummy wads so familiar to users of hard water.
This reaction consumes the soap, so that more must be
added to get a cleansing effect.

There are several ways of softening water. If the
water has only temporary hardness, it may be removed

by boiling (p. 102). As this requires fuel, a cheaper method consists in adding washing soda. It reacts as follows:

$$CaSO_4 \quad + \quad Na_2CO_3 \quad = \quad CaCO_3 \quad + \quad Na_2SO_4$$

| calcium sulfate | sodium carbonate | calcium carbonate | sodium sulfate |
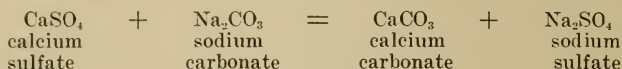
The calcium carbonate being insoluble in water settles to the bottom. The sodium sulfate does no harm. There is thus nothing left in the water to react with the soap. Before soap is added to hard water for laundering purposes, the water should be softened by the addition of washing soda. Much less soap will then be required to produce a suds. On a large scale water is softened in tanks by adding washing soda, allowing the precipitated calcium carbonate to settle to the bottom, and then drawing off the clear, soft water; or, in the case of large city water works, the sediment is removed by filtration.

QUESTIONS

1. How does soap act in removing dirt from cloth or from the skin?
2. Does soapy water dissolve fats?
3. What is meant by an emulsion? Is milk an emulsion? Why?
4. State two ways in which a grass stain might be removed from cloth.
5. What kinds of spots are best removed by gasoline, carbona, and chloroform? By water? By turpentine? By bleaching powder?
6. What is a soap? What is the source of glycerin?
7. Why are " fillers " added to soap?
8. Why are perfumes added to soap?
9. Why do the leachings from wood ashes make a soft soap?
10. Explain the action of hard water on soap. How can this be prevented?

LABORATORY EXPERIMENTS

100. **To Remove Stains by Bleaching.**—Make stains on pieces of cotton cloth with green grass, grape juice, red ink, and ordinary writing ink. Test the effectiveness of warm oxalic acid solution, hydrogen peroxide, bleaching powder solution, and alcohol in removing them. Compare the bleaching action of sulfur dioxide in experiment 46, and of chlorine in experiment 49.

101. **To Test the Effect of Alkalies and of Acids on Various Colored Fabrics.**—Collect a number of samples of colored goods of both cotton and silk and test the decolorizing effect of a drop of 5 per cent acetic acid and of a drop of 5 per cent of ammonia on each. Also test the restoring action of the acid on color removed by the alkali, and *vice versa*. Record the color of each piece of goods before treatment, after treatment with acid, and after treatment with ammonia.

APPENDIX

LIST OF REFERENCES TO SUPPLEMENT THIS TEXT.

Descriptive Chemistry. Newell. D. C. Heath & Co., Boston.
Chemistry of Common Things. Brownlee, Fuller, Hancock and Whitsit. Allyn & Bacon, Boston.
The Living Plant. Ganong. Henry Holt & Co., New York.
Productive Soils. Weir. Lippincott, Philadelphia.
Barnyard Manure. Beal. Farmers' Bulletin 192, U. S. Department of Agriculture, Washington, D. C.
Feeds and Feeding. Henry and Morrison. Published by the Authors, Madison, Wis.
Productive Feeding of Farm Animals. Woll. Lippincott, Philadelphia.
Principles of Human Nutrition. Jordan. Macmillan, New York.
Function and Uses of Food. Langworthy. Circular 46, U. S. Department of Agriculture, Office of Exp. Sta.
Successful Canning and Preserving. Powell. Lippincott, Philadelphia.
Bread and Bread Making. Atwater. Farmers' Bulletin 389, U. S. Department of Agriculture, Washington, D. C.
Some Common Disinfectants. Dorset. Farmers' Bulletin 345, U. S. Department of Agriculture, Washington, D. C.
The Chemistry of Cooking and Cleaning. Richards and Elliott. Home Science Pub. Co., Boston.
Home and Community Hygiene. Broadhurst. Lippincott, Philadelphia.
Dairy Chemistry. Snyder. Macmillan, New York.

LIST OF APPARATUS REQUIRED

For each ten students

1 balance, sensitive to 0.01 g.
1 set of weights, from 100 g. to 0.01 g.
1 hot air oven, about 10 x 10 x 12 in.
1 hot water oven, about 12 x 12 x 14 in. inside
1 500 c.c. cylinder
2 porcelain or glass mortar and pestles.
1 set cork borers
  Several dozen assorted rubber and cork stoppers
1 magnet
2 pneumatic troughs
4 files, round and triangular
4 wing-top burners
20 feet rubber tubing, 5 mm. inside diameter
30 feet glass tubing, 5 mm. inside diameter
4 thermometers, 0° to 120° C.
1 condenser (Liebig)
2 bell jars, 8 in. diameter, open at top
1 apparatus for electrolysis of water (optional)
4 florence flasks, 1000 c.c.
4 deflagrating spoons
10 feet aluminum wire, 2 mm.
½ pound copper wire, No. 20

For each student

1 Bunsen burner and tubing, or alcohol-lamp
1 ring stand and rings
1 filter arm
1 test tube holder
1 test tube brush
1 test tube clamp
1 test tube rack
1 asbestos wire gauge
1 pair tongs, 9 inches
25 filter papers, 9 cm.
2 porcelain evaporating dishes, 3 in.
12 test tubes, 6 in.
4 bottles, wide mouth, 8 oz.
4 glass plates, 4 in.
4 beakers, 400 c.c.
2 beakers, 100 c.c.
1 florence flask, 400 c.c.
1 funnel, 2 in., long stem
1 funnel, 2 in., short stem
3 watch glasses, 4 in.
2 stirring rods, 6 in.
1 box matches
1 cylinder, 100 c.c.
1 cylinder, 25 c.c.

Chemicals required by class of ten   Grams

	Grams
Acid, acetic (50 per cent.)	200
hydrochloric (sp.gr. 1.2)	2000
nitric (sp.gr. 1.4)	1000
sulfuric (sp.gr. 1.8)	2000
tartaric	150
oxalic	20
butyric	25
salicylic	20
molybdic	25
Alum	15
Alcohol, grain (ethyl)	500
wood (methyl)	200
Ammonium hydroxide (sp.gr. 0.9)	1000
tartrate	10
chloride	200
Barium chloride	100
Bleaching powder	100
Carbon bisulfide	500
tetrachloride	100
Calcium chloride	400
carbonate	100
sulfate	300
phosphate (tri-)	300
(acid)	100
hydroxide	100

	Grams
Cochineal	10
Charcoal, stick	30
Copper sulfate	150
oxide	50
Dextrose	40
Ether	700
Formalin	200
Ferrous sulfate	100
Hydrogen peroxide	100
Iodine	20
Iron filings	500
wool	200
wire, picture	100
Litmus paper, blue and pink	...
Lactose (milk sugar)	50
Lead acetate	10
Maltose (malt sugar)	50
Methyl orange	10
Manganese dioxide	30
Magnesium ribbon	100
sulfate	100
chloride	100
Mercurous chloride	20
Mercuric oxide	100
Phenolphthalein	10
Phosphorus, yellow	50
Pepsin	10
Potassium chloride	100
nitrate	25
iodide	10
permanganate	100
hydrogen phosphate	50
chlorate	100
sulfate	100
oxalate	20
Rochelle salt	200
Silver nitrate	20
Sulfur	500
Sodium nitrate	500
sulfate	100
hydrogen phosphate	100
hydroxide	500
Tannin	20
Zinc (flakes or granulated)	500

The above list does not include various common commodities which do not have to be bought through a chemical supply house, such as common salt, vinegar, sugar, seeds, various fruits and vegetables, baking soda, quicklime, limestone, soap, yeast, kerosene, gasoline, eggs, tea, copper wire, sand, Mason jars, starch, flour, linseed oil, cottonseed oil, and milk.

THE METRIC SYSTEM

```
     10 millimeters    (mm.) = 1 centimeter  (cm.)
     10 centimeters          = 1 decimeter  (dm.)
    100 centimeters          = 1 meter  (m.)
1000 cubic centimeters (c.c.) = 1 liter  (l.)
  1 cubic centimeter of water = 1 gram  (g.)
   1000 grams                = 1 kilogram, or kilo (kg.)
```

CONVERSION OF METRIC INTO ORDINARY UNITS

```
1 yard  = 91.6 cm.          1 meter = 39.3 inches
1 inch  = 2.54 cm.          1 cm.   = 0.39 inch
1 quart = 1134 c.c.         1 liter = 0.88 quart
1 oz.   = 28.3 g.           1 g.    = 15.4 grains
1 lb.   = 453.6 g.          1 kilo  = 2.2 lbs.
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